

ARMY MEDICAL LIBRARY  
WASHINGTON

Founded 1836



ANNEX

Section

*Chem.*

Number

*72373*







ELEMENTS

OF

CHEMICAL PHILOSOPHY,

ON THE BASIS OF REID,

COMPRISING

THE RUDIMENTS OF THAT SCIENCE

AND

THE REQUISITE EXPERIMENTAL ILLUSTRATIONS,  
WITH PLATES AND DIAGRAMS.

---

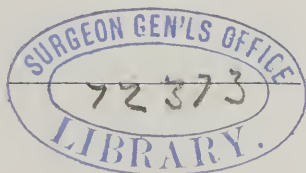
BY THOMAS D. MITCHELL, M. D.

Professor of Chemistry and Pharmacy in the Medical College of Ohio; President  
of the Ohio Medical Lyceum; Honorary member of the Philadelphia  
Medical and Columbian Chemical Societies, &c.

---

‘Plusieurs sciences, arts et metiers reposent presque uniquement sur la chimie, qui, de  
toutes les sciences, est celle dont on fait le plus d’applications aux besoins de la vie.’

*Berzelius.*



CINCINNATI:

PUBLISHED BY COREY & FAIRBANK.

NEW YORK: JONATHAN LEAVITT.

1832.

Entered according to act of Congress, in the year 1832,

BY THOMAS D. MITCHELL,

In the Clerk's Office of the District Court of Ohio.

---

A. F. ROBINSON, PRINTER.

To  
THE STUDENTS OF MEDICINE,  
AND  
LOVERS OF CHEMICAL SCIENCE,  
IN THE  
VALLEY OF THE MISSISSIPPI,  
THE FOLLOWING PAGES  
ARE  
RESPECTFULLY DEDICATED.

## ERRATA.

---

- Page 25, eighth line from the bottom; for *by the reason*, read *by reason*.  
30, fifth line from the top; for *affected*, read *effected*.  
30, seventh line from the bottom; for *effect*, read *affect*.  
35, fourth line from the bottom; for *forces of the solids*, read *pores of the solid*.  
132, in the table of the compounds of phosphorus with oxygen and hydrogen, the word *hydrogen* should be placed between the figures 16 and 1.  
175, to the title of section vii. add, *and its compounds*.  
188, twenty-fifth line from the top; for *as*, read *at*.  
251, first line; for *filled with acid gas*, read *filled with the acid gas*.  
435, fifth line from the bottom; for *are generally*, read *is*.  
454, thirty-seventh line from the top; for *cuuses*, read *causes*.  
461, sixth line from the bottom; for *residium*, read *residium*.

## P R E F A C E .

---

THIS book has been prepared, especially, for the use of the Medical Class in the College of Ohio. It is designed for the instruction of those who are desirous of learning, and has no pretensions to the character or claims of a copious and elaborate work.

To offer an additional treatise to the public notice, at a period when text-books are diffused through the land, may seem to be a work of supererogation. I have not, however, ventured on this enterprise, without due reflection, and flatter myself that the reasons which have prompted to the effort, are amply sufficient. Nearly all the books of chemistry which I have seen, are greatly defective in their arrangement, especially in their manifest departure from the fundamental principle of all instruction, viz. to pass, in regular gradation, 'from the known to the unknown.'

Moreover, in very many popular works, there is what I hold to be an unnatural division, and of course, a displacement of substances. Thus we read, in one part of a work, the history of metals, their uses, &c. and in a distant section, the salts of those metals are examined. Again, we find the properties and relative connexions of the gases discussed, and in some remote corner, a space is occupied with an account of the various modes by which those gases are prepared.

I had commenced the compilation of a book in alphabetical order, in the form of dictionary, to obviate, as much as practicable, the objections already stated, and many others, which lie against the most popular elementary books. While thus engaged, a new work, by Dr Reid, of Edinburgh, came into my possession; and finding it, after careful inspection, to accord almost wholly with the views I had long entertained, of the proper arrangement for a treatise on chemistry, I resolved to adopt it, as the basis of a text-book for my public lectures. Dr Reid's work commences with oxygen, and omits, altogether, the subjects of *affinity*, *caloric*, and *light*, as his design was to furnish a book of practical chemistry, rather than to give a full view of the science, in all its parts. Those deficiencies I have supplied, and in doing so, have borrowed freely from the *Cours de Chimie* of Laugier.

It will be seen, that the recent division of bodies into *electro-positive* and *electro-negative*, finds no place in this volume; and that some other modes of division, that have had considerable repute, are alike neglected. If this be indicative of defect, it is not accidental, but the result of design; because I am persuaded, that such divisions are not calculated to facilitate the progress of a student, in his first efforts to acquire a knowledge of the science.

In the arrangement of subjects which I have adopted, as laid down by Reid, the attentive observer will discover a progressive order, a gradual ascent, which is not to be seen in the ordinary treatises. Thus, after having examined the properties, formation, &c. of oxygen, the characteristic marks of hydrogen are noticed. Next follows water, a compound of oxygen and hydrogen, and a brief notice of the deutoxyde of hydrogen. Then comes nitrogen and all its compounds with oxygen and hydrogen. After this, sulphur, phosphorus, and carbon, together with the compounds resulting from the union of each with oxygen, hydrogen, and nitrogen, are brought into view. And as alcohol, ethers, the vegetable acids and alkalis, are compounds of carbon, hydrogen, and nitrogen, (elements already discussed,) these are examined in detail, as they can better be studied in this immediate connexion, than by assigning to them a remote place in the volume. There is a further advantage in this arrangement, and that is, the necessary attention which must be given to an important branch of vegetable chemistry, at an early period in the course of lectures. In almost all the schools, as well as in the books, vegetable chemistry comes in as a kind of rear-guard, and frequently, the lecturer dwells so long on other subjects, that no time remains for the examination of this interesting department.

In the present work, care has been taken to introduce all the practical applications of the science, that are important in medicine, and especially its toxicological bearings. I have added, also, a brief account of mineral waters, and the method of analysis; also, some remarks on combustion and nomenclature, which are not contained in Dr Reid's book. And, although a principal object has been to compile a suitable text-book for the convenience of medical students, I flatter myself, that men of other occupations will find the work to be calculated, in no small degree, to facilitate the study of a most delightful and important branch of natural science.

The period has at length arrived, when the intelligent members of our profession no longer doubt, that chemistry is essential to a good and thorough medical education. The deep interest manifested by a large proportion of the class in the Medical College of Ohio, at the last session, in this department of medicine, while it was gratifying to the lecturer, was also demonstrative of a happy advance in taste and good sense.

Let it be remembered, however, that the present performance is intended to be no more than an introduction to the science; and although, in this view, it is of sufficient extent for all the purposes of chemical study, during a course of lectures, it must be succeeded by the more diffuse and elaborate works of Henry, Silliman, Berzelius, and Thenard.

# EXPLANATION OF THE PLATES.

---

## PLATE I.

FIG. 1.—*A Glass Measure*, graduated into certain capacities, either by weight or bulk, to a standard fluid.

FIG. 2.—*The Balance or Pair of Scales*. This instrument is sufficiently known and duly appreciated. It should be kept in a dry place, in a glass case, so as to exclude the action of air and moisture. The most delicate balance is required for chemical experiments.

FIG. 3.—*Mortar and Pestle*. Pounding is a very frequent operation in the laboratory, and mortars of various kinds are required for substances of different textures. Glass, iron, marble, and Wedgewood are in general use.

FIG. 4.—*A Plain Retort*. Retorts are made of glass, earthenware, metals, &c. Glass retorts require some care, to prevent fracture. If any solid substance be put into them, which adheres to the bottom, a lamp heat will generally break them; the same result often follows, on laying a hot retort on a substance capable of conducting the heat from it suddenly. No danger need be apprehended from laying it on woollen cloth, or on dry glass, or on carded cotton or tow.

FIG. 5.—*A Plain Receiver*. Receivers are of various kinds and constructions. The glass vessels used over the pneumatic tub, to collect gases, are called receivers, but the one in the plate is such as is generally used in connexion with a retort.

FIG. 6.—*A Matrass*, is a glass vessel, with a large ball at one end, having a tube passing from it, of variable length. It is used, generally, for making solutions.

FIG. 7.—*A Cucurbite*. It is not much used at present. Retorts and receivers are more convenient.

FIG'S. 8 and 9.—*Spirit Lamps*. These may be constructed in many ways, and are always preferable to oil lamps, because the alcohol does not blacken the vessel exposed to the flame. The degree of heat is easily regulated, by elevating or depressing the wick.

FIG. 10.—*A Syphon*. This instrument is sometimes wanted in the laboratory, for drawing fluids out of vessels that do not admit of being moved. There are different forms of this instrument.

FIG. 11.—*The Common Funnel.* Funnels of glass, porcelain, tinned iron, &c. are useful in a laboratory; they should be of various bores and diameters.

FIG. 12.—*A Bolt-head.* This resembles the matrass, Fig. 6, only that it is much larger.

FIG. 13.—*Bladder with a Stop-cock.* This instrument is indispensable, and several of different sizes should be at hand.

FIG. 14.—*Bell Glass or Glass Receiver, with a Stop-cock.* These are very useful in transferring gases into other vessels. It is usual to have the stop-cock connected with the bladder (Fig. 13) of such a size as to screw on the stop-cock affixed to the receiver.

## PLATE II.

FIG. 1.—*The Pneumatic Trough or Cistern.* When used, the cistern *A*, is to be filled with water. A common tub will answer for ordinary purposes. A shelf is placed across the cistern, at about six inches from the top, and the water should cover the shelf, at least one inch; *a*, is a glass jar or receiver full of water, and inverted on the shelf, directly over a hole made in the latter, to allow the gas to pass up, as it issues from the retort; *b*, is a eudiometer or detonating tube. *B*, is a retort, fixed on a lamp stand, and containing the materials for making the gas. It is heated by the lamp directly under it.

FIG. 2.—*A Tubulated Retort.* It differs from the plain retort (Fig. 4, Plate I) in having an aperture in the upper part of its body, to which is accurately fitted, a glass stopper. This contrivance is often very convenient; it allows the introduction of additional or new materials, while an operation is going on, and without disturbing the process.

FIG. 3.—*A Differential Thermometer.*

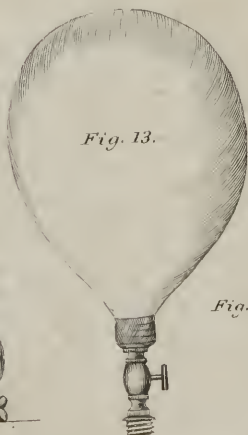
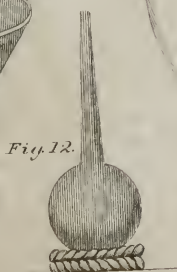
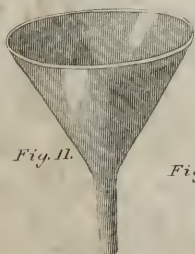
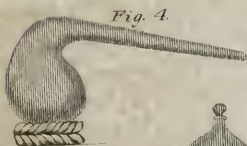
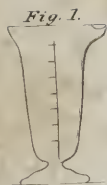
FIG. 4.—*Pictet's Apparatus for showing the reflection of heat.* *a c* and *b d*, are concave metallic mirrors, usually of polished tin or brass, fixed on stands; *f*, is a heated iron ball, placed directly in the focus of the mirror *b d*; *e*, is a thermometer, placed in the focus of the opposite mirror. The dotted lines show the course of the rays of heat diverging from the heated ball *f*, then rendered parallel by the mirror *b d*, whence they are reflected to the mirror *a c*, from which they converge and fall on the thermometer in the focus *e*.

FIG. 5.—*A Voltaic Pile*, being one of the earliest instruments to demonstrate galvanic electricity.

FIG. 6.—*A Galvanic Trough.* Various forms of this instrument are now in use. More commonly, however, the copper and zinc plates being soldered in pairs, are cemented in grooves at suitable distances from each other. *A A*, represent the positive and negative wires, passed into a glass tube, and nearly approximating, for the purpose of decomposing water.



Plate I.



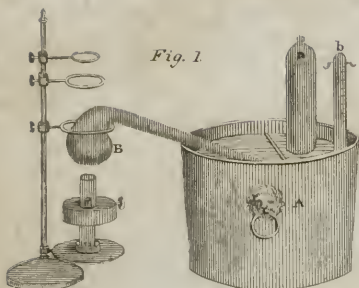


Fig. 1.

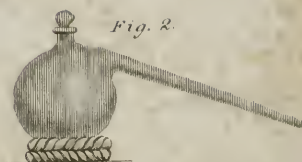


Fig. 2.

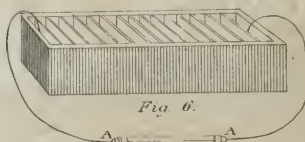


Fig. 6.

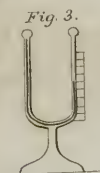


Fig. 3.

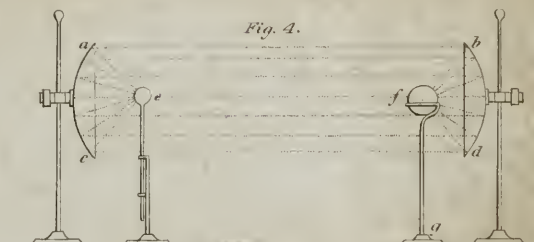


Fig. 4.



Fig. 9.

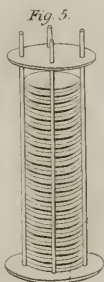


Fig. 5.

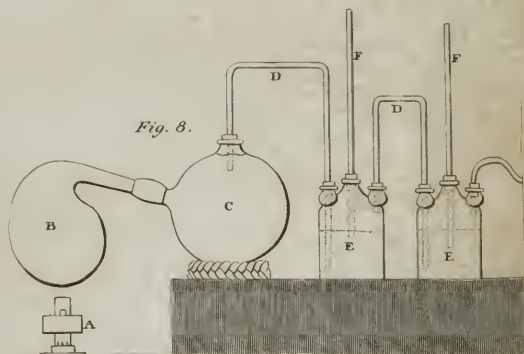


Fig. 8.



Fig. 11.



Fig. 7.



Fig. 12.



Fig. 10.

FIG. 7.—*A Glass Tube for the detonation of gases.* Near the upper end, there are two opposite perforations, through which platina wires are made to pass. These wires, on the outer side of the tube, are terminated by knobs.

FIG. 8.—*A Woulfe's Apparatus.* This is a useful instrument in various distillations. *A*, is a lamp to heat the retort *B*, the neck of which is accurately fitted and luted to the tubulated receiver *C*. Out of the tubulure of this receiver, passes the bent glass tube *D*, down nearly to the bottom of the Woulfe's bottle *E*; these bottles (Woulfe's) have, each, three apertures, all of which are important and indispensable. From the central aperture of *E*, the straight tube *F*, passes; its lower extremity dips an inch below the surface of the contained fluid, while the superior extremity is from five to ten inches above the neck of the bottle. The tubes *F F*, are called *safety tubes*, because they protect the apparatus from explosion. The second and third bottles *E E E*, (if three be used,) are arranged in the same manner as the first, already described, and from the last of the series, proceeds a bent tube, whose extremity dips into a vessel of water. Now, if by accident, the heat under the retort be unduly diminished, a vacuum will take place, cold water will be absorbed by the bent tube, and rushing back, will not only spoil or injure the contents of the Woulfe's bottles, but by its contracting power on the heated retort, it will cause that to be broken. This state of things is on the supposition that no *safety tubes* are employed. When these tubes are used, accident is either prevented or greatly diminished. When absorption of water takes place in the bent tube, the elastic gas in the adjoining bottle, instead of forcing the liquid into the next bottle, drives it up the safety tube.

FIG. 9.—*A Glass Bottle with Two Apertures*, from one of which passes a tube, gradually tapering to a point at the upper end. A Woulfe's bottle will answer equally well. The tube being closely fitted to one neck of the bottle, iron or zinc filings are passed into the bottle through the other aperture and afterwards diluted sulphuric acid is poured in. Hydrogen gas is thus formed, and on escaping by the tube, may be inflamed, and then constitutes the *philosophical candle*.

FIG. 10.—*A Glass Receiver, for burning fine iron wire in oxygen gas.*

FIG. 11.—*A Glass Globe with a Stop-cock, for weighing gases.*

FIG. 12.—*A Common Oil Flask, with a Bent Glass Tube.*

In addition to the articles noticed in the plates, *furnaces* may be mentioned. They may be constructed of brick or of sheet iron. A *chauffer* is also a very convenient utensil; it is simply a small furnace, made of sheet iron, having a few apertures near the bottom, which act as so many vents and maintain the combustion of the coal. It differs but little from a chafing dish.

Common sweet oil flasks, well cleansed by means of a solution of potash,

are very useful. If perforated corks be fitted to their necks, and bent tubes be luted to the apertures made in the corks, we have, at little expense, a very convenient piece of apparatus. The best mode of perforating corks for this use, is to have iron rods made quite round, and varying in thickness from one to three-eighths of an inch, and having heated them to redness, pass them through the centre of the cork. A clean and uniform hole is thus made, and by having a tube of proper size, and coating it with beeswax where you wish to insert it into the cork, you can easily succeed in making an instrument that will prove exceedingly useful.

Wine glasses, jelly glasses, tumblers, and the like, are also important, and with a small expense, the novice in chemistry may supply himself with all the articles that are requisite in the performance of operations which he ought to attempt.

# CONTENTS.

---

## PART I.

### CHAPTER I.

	Page
General Remarks ; Division of Bodies ; Various kinds of Attraction ; Phenomena of Combination ; Influence of external Agents on Decomposition, - - - - -	1 to 18

### CHAPTER II.

<i>Of Imponderable Bodies</i> , viz. Light ; Caloric ; Electricity ; Galvanism ; Magnetism, - - - - -	18 to 41
---	----------

### CHAPTER III.

<i>Of Nomenclature</i> , - - - - -	41 to 46
------------------------------------	----------

### CHAPTER IV.

<i>Of Specific Gravity</i> , - - - - -	46 to 49
--	----------

---

## PART II.

OF SIMPLE OR UNDECOMPOSED SUBSTANCES, - - - - -	50
---	----

### CLASS I.

SIMPLE SUBSTANCES, NOT METALLIC, AND THEIR COMBINATIONS WITH EACH OTHER.

#### CHAPTER I.

<i>Of Oxygen</i> , - - - - -	50 to 60
------------------------------	----------

#### CHAPTER II.

<i>Of Hydrogen</i> , and its Compounds with Oxygen, viz. Water and Deutoxyde of Hydrogen, - - - - -	60 to 76
---	----------

#### CHAPTER III.

<i>Of Nitrogen</i> , and its Compounds with Oxygen and Hydrogen, viz. Atmospheric Air ; Protoxyde of Nitrogen ; Deutoxyde of Nitrogen ; Hyponitrous Acid ; Nitrous Acid ; Nitric Acid ; Ammonia - - - - -	76 to 103
---	-----------

## CHAPTER II.

Page

<i>Of Barium and Strontium</i> ; Oxydes of Barium ; Protoxyde and Deutoxyde ; Barytes ; Hydrate of Barytes ; Sulphate of Barytes ; Sulphuret of Barium ; Hydrosulphuret of Barytes ; Nitrate of Barytes ; Carbonate of Barytes ; Muriate of Barytes ; Chloride of Barium ; Chlorate of Barytes ; Strontia or Strontites ; Salts of Strontites, - - - - -	351 to 356
--	------------

## CHAPTER III.

<i>Of Magnesium</i> ; Discovered by Tondi and Ruprecht, in 1790 ; Oxyde of Magnesium ; Magnesia ; Muriate of Magnesia ; Sulphate of Magnesia ; Incompatibles ; Phosphate of Ammonia and Magnesia ; Carbonate of Magnesia ; Bi-carbonate ; Liquid Magnesia, - - - - -	356 to 361
--	------------

## ORDER III.

OF GEOFIABLE METALS, - - - - -	362
--------------------------------	-----

## CHAPTER I.

<i>Of Aluminum</i> ; Oxyde of ; Alumine ; Alum ; Varieties of ; Incompatibles ; Alumen Ustum ; Homberg's Pyrophorus ; Acetate of Alumine, - - - - -	362 to 369
---	------------

## CHAPTER II.

<i>Of Silicium</i> ; Oxyde of Silicium or Silica ; Silicates, -	369 to 372
---	------------

## CHAPTER III.

<i>Of Glucinium, Yttrium, Zirconium, and Thorinum</i> ; Glucina, Yttria, Zirconia, and Thorina, - - - - -	372
---	-----

## ORDER IV.

OF METALS WHOSE OXYDES CANNOT BE REDUCED BY HEAT ALONE, - - - - -	372
---	-----

## CHAPTER I.

<i>Of Iron</i> ; Oxydes of ; Sulphurets of Iron ; Iodide of Iron ; Phosphuret ; Carburets ; Muriate ; Muriated Tincture ; Chloride of Iron ; Nitrate of Iron ; Sulphate of Iron ; Incompatibles ; Carbonate ; Phosphate ; Acetate ; Tartrate ; Gallate ; Ferrocyanate, - - - - -	373 to 387
--	------------

## CHAPTER II.

<i>Of Lead</i> ; Its Action on Water ; Oxydes ; Chloride ; Iodide ; Sulphuret ; Salts of Lead ; Nitrate ; Sulphate ; Phosphate ; Carbonate ; Acetate ; Incompatibles ; Extract of Lead, -	387 to 397
---	------------

## CHAPTER III.

<i>Of Copper</i> ; Oxydes ; Chlorides ; Sulphurets ; Sulphate ; Bisulphate ; Incompatibles ; Sulphate of Copper and Ammonia ; Carbonate of Copper ; Nitrate of Copper ; Bi-nitrate ; Acetate ; Bi-acetate ; Subacetate ; Protomuriate ; Permuriate ; Poison of Copper ; Tests, - - - - -	397 to 406
--	------------

## CHAPTER IV.

	Page
<i>Of Zinc ; Spelter ; Granulated zinc ; Oxyde ; Chloride ; Sulphuret ; Sulphate ; Incompatibles ; Carbonate ; Nitrate ; Acetate ; Muriate,</i>	406 to 411

## CHAPTER V.

<i>Of Antimony ; Crude ; Regulus ; Oxydes ; James's Powder ; Chlorides ; Sulphurets ; Sulphate ; Tartrate of Antimony and Potash ; Incompatibles ; Muriate,</i>	411 to 428
---	------------

## CHAPTER VI.

<i>Of Arsenic ; Oxydes ; Arsenic and Arsenious Acids ; Arsenites and Arseniates ; Tests ; Substitute for Arsenic to kill Rats ; Sulphurets ; Chloride ; Antidotes for poison of Arsenic,</i>	428 to 443
--	------------

## CHAPTER VII.

<i>Of Tin ; Tin Stone ; Block-tin ; Granulated tin ; Tin-foil ; Tin Filings ; Oxydes ; Chlorides ; Sulphurets ; Muriate,</i>	443 to 446
--	------------

## CHAPTER VIII.

<i>Of Bismuth ; Oxyde ; Chloride ; Sulphuret ; Nitrate ; Magistery of ; Tartrate ; Pearl white ; Flake white,</i>	446 to 449
---	------------

## CHAPTER IX.

<i>Of Manganese ; Oxydes ; Manganetic and Manganeseous acids ; Chameleon Mineral ; Chloride ; Fluoride ; Sulphuret,</i>	449 to 456
---	------------

## CHAPTER X.

<i>Of Chrome ; Oxydes ; Chromic acid ; Fluochromic acid ; Chlorochromic acid ; Salts of Chrome ; Chromate and Bi-chromate of Potash ; Chromate of Barytes ; Chromate of Lead ; Chromate of Mercury,</i>	456 to 461
---	------------

## CHAPTER XI.

<i>Of Cobalt ; Smalt ; Zaffre ; Powder blue ; Oxydes ; Muriate ; Sympathetic Ink,</i>	461 to 463
---	------------

## CHAPTER XII.

<i>Of Tungsten ; Oxydes ; Tungstic acid ; Tungstates,</i>	463 to 464
---	------------

## CHAPTER XIII.

<i>Of Molybdenum ; Oxydes ; Molybdic acid ; Molybdates,</i>	464 to 465
---	------------

## CHAPTER XIV.

<i>Of Uranium ; Uranite ; Oxydes,</i>	465
---------------------------------------	-----

## CHAPTER XV.

<i>Of Columbium ; Tantalum ; Oxyde ; Columbic acid ; Chloride of Columbium ; Sulphuret of Columbium,</i>	465 to 466
--	------------

## CHAPTER XVI.

<i>Of Titanium ; Oxydes ; Titanic acid ; Cerium ; Cadmium ; Tellurium ; Pluranium ; Vanadium,</i>	466 to 467
---	------------



## ORDER V.

Page

OF METALS WHOSE OXYDES CAN BE REDUCED BY EX-  
POSURE TO HEAT, WITHOUT THE AID OF  
INFLAMMABLE MATTER.

## CHAPTER I.

*Of Mercury* ; Its various Names ; How purified ; Oxydes ;  
Æthiops per se ; Blue pill Mass ; Tests of Mercury in solution ;  
Salts ; Chlorides ; Iodides ; Sulphurets ; Incompatibles ; Anti-  
dotes, - - - - - 467 to 493

## CHAPTER II.

*Of Silver* ; How obtained pure ; Cupellation ; Oxyde ; Salts ;  
Chloride ; Iodide ; Cyanide ; Sulphuret ; Tests, - - - 493 to 500

## CHAPTER III.

*Of Gold* ; Origin of the Word ; General Properties ; Quartation ;  
Parting ; Oxydes ; Chloride ; Ammoniuret ; Ethereal or Pota-  
ble Gold, - - - - - 500 to 503

## CHAPTER IV.

*Of Platina* ; Spongy ; Oxydes ; Sulphate ; Fulminating Platina ;  
Chloride used as a Test, - - - - - 503 to 505

## CHAPTER V.

*Of Nickel* ; Copper-nickel ; Speiss ; Oxydes, - - - 505 to 506

## CHAPTER VI.

*Of Palladium* ; Rhodium ; Osmium ; Iridium ; Found in Ores of ;  
Platina, - - - - - 506 to 507

ALLOYS AND AMALGAMS, - - - - - 507 to 509

## PART III.

*Of Animal Substances* ; Fibrine ; Albumen ; Gelatine ; Osma-  
zone, - - - - - 510 to 513

*Of Miscellaneous Experiments*, illustrating the Examination of  
Calculi, - - - - - 513 to 516

## PART IV.

*Of Animal and Vegetable Analysis* ; Analysis of Mineral Waters ;  
Of Coloring Matters, Tests, &c. ; Lutes and Cements ; Elec-  
tricity and Galvanism ; Acidimetry and Alkalimetry ; Tables of  
Weights and Measures ; Correspondence of Thermometers ;  
Freezing Mixtures, - - - - - 517 to 554



# ELEMENTS

OF

## CHEMICAL PHILOSOPHY.

---

### PART I.

---

#### CHAPTER I.

General Remarks. Division of Bodies. Various Kinds of Attraction. Phenomena of Combination. Influence of external Agents on Decomposition. Application of the Laws of Chemical Affinity.

THE term philosophy, embraces a variety of important sciences, each of which has modes of illustration, peculiar to itself. Thus, for example, geometry measures the angles and the various dimensions of bodies; the naturalist takes cognisance of the external forms of bodies, of their color, odor, structure and all their obvious qualities; while the chemist, not unmindful of these points of study, superadds modes that are perfectly *sui generis*. He looks into the minute particles or atoms of substances, penetrates into their very centre, detaches their ingrant parts, and endeavors to determine the elements which make up their composition. Hence, arose a definition of chemistry, that was formerly very popular, viz. that it had for its object, the knowledge of the molecular and mutual action of bodies on each other.

Such has been the progress of chemical science, that it now takes a more extensive range; and just in proportion to the multiplication of discoveries, will be the magnitude of its scope and influence.

A striking difference between natural philosophy and chemistry, consists in the nature of the subjects over which they preside. The former takes cognisance of matter in mass, while the latter acts upon the imperceptibly minute atoms of bodies.

Hence we have the ancient adage, *Ubi desinit physicus, incipit chemicus*.

In the view of the chemist, all bodies in nature, are either *simple* or *compound*. By the former, he means all those bodies which have, thus far, resisted the most laborious efforts at decomposition; such as have never yielded more than one substance or kind of matter, to the most untiring analyst. Hence, we are warranted in designating as *simple*, many bodies, which, in the progress of science, may yet prove to be compounds; and until their mixed character be shown as the result of actual experiment, we are bound by the very nature of the case, to style them *simple*. Thus, there was a period, when this appellation was proper, in reference to *air, fire, earth, and water*; for prior to the days of Scheele and Cavendish, there were no certain data, by virtue of which, these substances could be excluded from the list of simple bodies. They were, moreover, the only simple substances that chemists recognised, all others being, in their estimation, compounds. How great the revolution has been in this respect, may be inferred from the fact, that the four original elements are all compounds, while more than fifty simple bodies are now enrolled on the pages of our chemical treatises.

A simple substance is, therefore, made up of only one sort of matter, while, at least, two varieties are essential to every compound; thus, oxygen and hydrogen are simple, while water is a *compound*, because it contains both of these simple substances.

In order to determine whether a body is really simple or compound, something more than mechanical force must be employed; for this, alone, can never separate the heterogeneous and ultimate particles of matter. We may, indeed, divide a body, and subdivide it into very small parts, by means of a hammer or a pestle and mortar, or filing, or other mechanical agency; but while these forces reduce the mass to particles of almost inconceivable minuteness, they do not move a hair's breadth in the work of decomposition. Each fragment, though fine as the dust in the balance, retains the identical qualities of the primitive mass. Thus, if we divide a piece of loaf sugar, *ad infinitum*, each grain will exhibit the peculiar properties of sugar, as perfectly as did the original loaf.

A piece of pure iron or lead, is composed of simple homogeneous particles, because these bodies are, in themselves, simple. When, by mechanical means, we detach their particles from each other, we of course separate simple particles from others equally simple, and previously held together by simple attraction or cohesion. We cannot act thus, however, on a lump of cinnabar, because it is a compound, made up of two

simple bodies. We may, indeed, fracture the lump, and so obtain from it many smaller pieces of cinnabar; but each distinct piece, however small, is a compound, like the original lump, of sulphur and mercury; and although these are altogether dissimilar in their properties, it is impossible to break up their connexion by any mechanical power that can be applied.

The term *integrant particles*, has been given to those which are procured by mechanical division, and which are, of consequence, similar in all respects, to the mass of which they were a part. Hence, if they be taken from *simple* bodies, as iron, these *integrant particles* will also be simple; if detached from *compound* bodies, they will necessarily be compound. It is, therefore, perfectly correct to speak of the integrant particles of iron, and of sugar, or of cinnabar. Iron filings are consequently simple integrant, while grains of sugar are compound integrant, particles. If we take some of the latter, and expose them to the influence of chemical agents, we produce a change, which mechanical effort is inadequate to accomplish, viz. we resolve them into their *constituent* elements or particles. So also, if we divide a bason of water, into as many drops as is possible, each drop will continue to be water, and may be called a compound integrant particle, because water is a compound. But, if a portion of this liquid be acted on by galvanism, or other means, to be hereafter detailed, and which are purely chemical, these compound integrant particles will be resolved into hydrogen and oxygen, which are designated the *constituent* particles of water. The cause, in virtue of which the particles of matter (whether *compound* or *simple* integrant) adhere, has been a source of speculation, in all ages, and yet, even in these days of improvement, we know nothing that is satisfactory on this point. Some have regarded it, as analogous to, or synonymous with, *planetary* attraction; they have supposed that the power which retains the stars in their orbits, was equally operative in the cohesion of common matter. And when this subject is viewed in reference to the great first cause, it is abundantly evident, that the proposition is strictly true. It is, in a word, the same power exerting itself in the one case at great distances, and between very large masses, that operates in the other, on molecular atoms, and at distances of inconceivable minuteness.

Let this point be settled as it may, it is known, that different names have been given to the power by which particles are held together; that which operates upon atoms of the same nature, is the *force of cohesion*, or *attraction of aggregation*. It holds all masses together, whether they be exclusively simple or actually compound in their character. Of course, it is the bond

of union, between *simple* as well as *compound* integrant particles; for a separation of those particles can always be accomplished by mechanical agency. Diverse from this power, is the force or energy employed to hold together the *constituent* particles of bodies, since these are wholly undisturbed by mechanical action. To this force, we give the appellation of *chemical affinity*, *chemical attraction*, and *attraction of composition*. We say, therefore, that the particles of a simple substance *cohere*, or are bound in mass by the attraction of cohesion; and we affirm that when two or more dissimilar particles unite to form a compound body, the result is the effect of *chemical attraction*, or the attraction of composition.

If these points be well established, the inference is unavoidable, that the reciprocal action of similar or homogeneous particles on each other, occasions an aggregation or a mass; and this mass is, in no respect, different from the particles out of which it has been formed, except in bulk.

Exp. A tube of glass is plunged into water, and when drawn out, a drop is visible at the inferior end, which does not fall from the tube. Now it is plain, that some cause operates to hold the drop in that situation, and that cause is simply *cohesion*. I take an extremely fine powder, and by the aid of a

Exp. very little aqueous matter, the whole is clotted together, and a mass is the result; but why? I answer, because the particles are joined by cohesion. Take two panes of glass

Exp. and apply them to each other, having first interposed a little water. They are found to adhere; the inferior pane does not fall, and is really capable of holding a considerable weight. Now this phenomenon cannot be explained satisfactorily, except by reference to the energetic power of cohesion. This same power causes the solid particles that are suspended in a liquid, to reunite when precipitated by the addition of a suitable agent. Thus camphor, which is a solid, may be dissolved in alcohol, and then its solidity is lost. If water

Exp. be added to this solution, it combines with the alcohol and the camphor resumes its solid form. In like manner, when crystallisable substances are held in solution, evaporation gradually dissipates the solvent, and in exact proportion to this effect, is the tendency of the crystallisable matter to cohere, and regular crystallisation follows.

Chemical attraction and the attraction of cohesion are antagonist forces; they not only stand opposed to each other, but they always bring about different results. It is impossible to excite chemical action, vigorously, until the cohesive power be destroyed. So long as similar particles remain in contact, they

will not be disposed to combine with particles of another sort. Thus, I pour on a *solid* mass of metallic copper, the strongest nitric acid, and, in another vessel, I place the *filings* of copper and a quantity of the same acid. The difference of action in the two cases is instantly seen, and may be easily understood. In the first instance, the solid mass is presented, and before the acid can effect combination, it must gradually detach the integrant particles. In the latter, the particles are already separated and the process is rapid and tumultuous. Exp.

It is, therefore, to be received as a law of chemical action, that a prerequisite is the destruction of the attraction of cohesion. Hence, before we attempt to bring a solid into a state of chemical combination, we resort to the expedient of pulverising, and thus, by minutely dividing the solid, we increase, almost beyond calculation, the points on which chemical affinity may exert its energies. Hence, the saying of the ancients, *Corpora non agunt nisi sint fluida*, was not so much opposed to fact, as some have imagined. If the idea of fluidity had reference only to the minute division of bodies, the axiom might be admissible in the present times; we have several cases which show clearly, that chemical action will not take place, unless one of the agents be in the fluid state. The solvent employed, acts only by separating still further the solid particles, and thereby augmenting the points of contact, at which chemical action is energetic. Thus, powdered tartaric acid and bi-carbonate of soda, may be mixed and triturated in a mortar, and no change be effected. If water be poured on, or if the tartaric acid be dissolved in water, and then added to the bi-carbonate, effervescence speedily takes place and a new compound is formed. Exp.

There are some instances in which two solid substances may enter into chemical combination; but in all these, I believe the effect to be produced by the extreme division of the particles, as the result of active trituration; of such bodies it may be correctly affirmed, that, although in state of powder, the attraction of cohesion was insufficiently destroyed, to permit the operation of chemical affinity, until the pestle and mortar had more completely effected the division of the particles. Thus, if quick lime and sal-ammoniac be thrown into a mortar, in the state of powder, and gently stirred with a stick, no change is perceptible. Apply the pestle vigorously to this mixture, and in a few minutes, a most pungent odor is evolved, and you are certain, that a chemical change has occurred. This odor comes from the sal-ammoniac; it is ammoniacal gas, which is a constituent part of that compound. That this odoriferous Exp.



gas has been separated by the chemical agency of the quick lime, and not by mere trituration, is plain from this circumstance; you may rub the sal-ammoniac, alone, in a mortar, for a month, and you will look in vain for a similar evolution.

We, moreover, infer that cohesion is the great antagonist of chemical action from the readiness with which liquids and gases enter into combinations; in these, the cohesive power is almost lost. Indeed, when we cause gases to combine, we are unable to perceive any thing like an obvious disruption of the force of cohesion, because that force is too feeble to be estimated.

In like manner, we effect, without difficulty, the union of a solid and a liquid; thus common salt and sugar dissolve in, and readily combine with, water; and, under favorable circumstances, we combine a solid or a liquid with a gas. So if iron or mercury be exposed to the twofold action of a powerful heat and atmospheric air, the metal combines with oxygen, and the product is what chemists call an oxyde. Now the heat overcomes, by its repulsive force, the cohesion of the particles of metal, and thus enables the air to exert a chemical action on those particles, which it could not effect so long as the cohesive force predominated.

I have already intimated, that the combination, resulting from chemical affinity, is more energetic, in proportion as the points of contact are increased; it may be added, further, that such combination never occurs, but at the points of contact. This position is susceptible of illustration, by a very simple experiment.

Exp. Introduce into a glass tube, an aqueous solution of litmus, to which a small quantity of common salt has been added, to augment its density. Pour in gradually some acetic acid, whose density has been diminished by solution in alcohol. At the moment when these liquids come in contact, a red line is perceptible, resulting from combination. If the tube be now shaken, so as to mix the liquids, and thus multiply the points of contact in all directions, the whole mass will assume a red color, because the union is now perfect or entire.

Whenever particles of matter enter into combination, by the force of chemical attraction, we perceive, either at the moment, or soon afterwards, several important phenomena, which are never observed in the operations of the attraction of cohesion; very generally, there is a sensible development of heat, and, occasionally, we notice a reduction of temperature. In some cases, the heat evolved is scarcely perceptible, while in others, it is intense, and this diversity depends on the rapidity with which bodies combine. If the union be gradually effected,

the temperature will be augmented in a very feeble degree; if combination result by an instantaneous movement, the evolution of heat will be among the most obvious phenomena accompanying the experiment.

That a great deal of heat is given out, in the process called *slaking of lime*, is well known to every one. The water poured on the lime disappears, and the solid, instead of becoming fluid, is simply reduced to a pulverulent condition. The union, effected between the water and the lime, is quick, though imperceptible, and the heat evolved, is correspondingly great. In manufacturing the article, called *precipitate per se*, the mercury is exposed to heat and atmospheric air, for about fifteen days; and, although a positive chemical union is effected between the metal, and one of the constituents of the air, yet the length of time necessary for the operation, is so great, as to prevent us from recognising the fact, that heat is constantly evolved, from the first to the last day of the process.

If four parts of the article usually called oil of vitriol be poured on one of water, there will be so great an evolution of heat, as to cause the fracture of a glass vessel, if the mixture be made in it. Here, combination is almost instantaneous, and the heat is proportionably great. Let the same quantity of oil of vitriol be poured into fifty times as much water, and there will be very little augmentation of temperature, because the combination is comparatively slow. Exp.

Again, let a piece of phosphorus, half the size of a pea, be placed in a saucer and exposed to the open air; we perceive in a short time, by the vapor around the phosphorus, that some sort of action is going on. It is a slow combustion, and although the whole of the solid would in a given time be consumed, yet very little, if any, heat would be evolved. Take another piece of the same article, of similar dimensions, and place it on another saucer, and apply a lighted taper or a red hot wire. The whole is, at once, in a blaze, and the solid is changed into dense vapors, and we are conscious that a very considerable amount of heat has been given out. But, if another piece of the same solid be inflamed and dipped quickly into a vessel of oxygen gas, the combustion is much more rapid, and the elevation of temperature, decidedly more conspicuous. We have here, three grades of intensity of chemical action, as exhibited by the same solid substance; and the diversity in the evolution of heat, results entirely from the difference in these grades. Exp.

Connected with the evolution of heat, which attends chemical action, we notice a contraction or condensation of the sub-

stances that have been combined, though this is not always the case. If we pour four or five parts of oil of vitriol, into a glass tube of a bore less than half an inch, and add one or two parts of water, and mark the height in the tube to which the fluids

rise, we shall discover, on shaking the tube, and then setting it aside to cool, that the mixture occupies a less space than before. Now it is impossible to explain the considerable reduction of bulk, without a reference to actual condensation; and we cannot solve the problem, unless we admit that the particles of both the fluids employed, although united to each other, were separated by inappreciable distances, which having been overcome by the force of chemical action, induced a still closer approximation of atoms, and so diminished the actual volume. This notion of interstices between the particles of matter, is further supported by the well known fact, of the shrinking or contraction of the mercury in a thermometer, removed from a warm room, to the open, wintery blast.

The operations of chemical affinity are accompanied by other phenomena, which are not less important. Thus, we notice a change of *condition*, of *color*, of *taste*, of *smell*, and of the *action on the animal economy*. Two gases for example, by mixture, enter into chemical combination; the gases disappear and a solid is the result. This is seen, when ammoniacal gas, and muriatic

acid gas are passed, by different apertures, into a common receiver. Dense, white clouds appear, and solid muriate of ammonia is the product.

Into a glass vessel, partly filled with a transparent solution of muriate of lime, pour a small quantity of colorless oil of vitriol. No sooner do these liquids come in contact, than a solid mass is formed.

On the other hand, we sometimes observe the change from solidity to fluidity, and even to the gaseous state; what is called the *carburet of sulphur* is a liquid, and yet it is formed by the union of two solids, viz. charcoal and sulphur. So also, the sugar of lead and the white vitriol of the shops, are both solids, but if we triturate them in a glass or marble mortar, chemical action ensues, and the mass becomes fluid.

There are also changes in *color*, which are the result of chemical action or affinity. Thus, a colorless gas, uniting with mercury and chrome, both of which are nearly white, gives birth to

*red* and *green* compounds. If *colorless* solutions of prussiate of potash and sulphate of iron are mixed, a *deep blue* is produced. *Colorless* hydriodate of potash and corrosive sublimate in solution, yield a beautiful *scarlet red*; if the former be added to sugar of lead, which is



also without color, a *bright yellow* is formed. The *taste* and *peculiar properties* of bodies, are changed, by chemical action. Thus, oil of vitriol is intensely sour and corrosive; soda is very caustic and possessed of an alkaline or ley taste. But the salt which results from the chemical combination of these caustic substances, is totally different from either of its component parts. It is known by the name of *Glauber's salts*, and is familiar to almost every one, for its mild, cathartic properties.

Sulphur and the gas called hydrogen, when perfectly pure, are destitute of smell, but when combined chemically, in the form of what is denominated sulphuretted hydrogen, we have a most offensive and disgusting article. So also nitrogen and hydrogen, in a state of purity, emit no odor, but ammonia, which is a compound of those two gases, is exceedingly pungent. The atmospheric air which we are continually inhaling, is a mixture, chiefly of oxygen and nitrogen, in certain proportions. But if the same substances be combined in other proportions, say five of the latter to one of the former, we have a compound, which is not only irrespirable, but absolutely deleterious, in the highest degree.

Hence, the truth of a statement already made, will be properly appreciated at this stage of our subject, viz. that the attraction of composition or chemical affinity, operates on matter, in a way that is *sui generis*, and totally different from the action of the forces of cohesion; while the latter fails entirely to change the intrinsic properties of matter, the former never leaves bodies in the precise state in which it found them, but invariably alters, and sometimes wholly destroys, the original properties.

Some substances are susceptible of combination, in all proportions, as water and oil of vitriol, or water and spirit of wine; but in these cases, the power of affinity is, comparatively, weak. Bodies, between which there is a very strong affinity, do not combine in this unlimited manner, but always in fixed and definite proportions. Hydrogen and oxygen combine in only two varieties, viz. water and the deutoxyde of hydrogen. In ammonia, we have the only compound, which can possibly result from the union of hydrogen and nitrogen.

For the most part, bodies unite in *two*, *three*, sometimes *four*, and very rarely, in *five* proportions. The latter are the most complex substances, with which chemists are acquainted, and fortunately, are seldom seen. A striking feature in this view of combination, is, that among the different uniting proportions of the same bodies, there is an intimate and uniform relation, whereby the quantity of one of the two bodies which enter into a second combination, is a multiple, by a whole number,

of the quantity contained in the first combination. For example, *A* combines with *B* in fixed proportions, and gives rise to what is called a *binary* compound. If these two substances may be combined in a second proportion, it will appear, that although *A* remains stationary, *B* will be doubled, or perhaps increased threefold. Let us suppose that in the first compound of *A* and *B*, they are as 1 to 1; in the second, *A* will continue to be 1, and *B* will be increased to 2; in the third compound of *A* and *B*, the former continues unchanged, while the latter is as 3. Here it is plain, that *B* in the second and third combinations, is augmented above its first compound, by a multiple of a whole number. The amount 1 of *B*, in the first compound, is exactly doubled and trebled, in the two subsequent combinations. We observe this relation in the metallic oxydes, as well as in other compounds. Thus mercury is susceptible of two combinations with oxygen, in the first of which 200 of mercury are combined with 8 of oxygen, constituting the *black* oxyde; we have also the red oxyde, composed of 200 mercury, and 16 oxygen. There are also two sulphurets of mercury, in which the same kind of relation obtains. In the first, 200 hundred of mercury are joined to 16 of sulphur; and in the second, the same quantity of metal is combined with 32 of sulphur. So in all the acknowledged chemical compounds of nitrogen and oxygen, we observe the same law. The proportion of nitrogen remains unaltered, while the quantity of oxygen is not the same in any two of the compounds. They vary thus:

Nitrogen.	Oxygen.
14 - - - - -	8
14 - - - - -	16
14 - - - - -	24
14 - - - - -	32
14 - - - - -	40

In all this variety of combination of two bodies, we discover a regular gradation in the distribution of one of the constituents, while the other experiences no change. Occasionally, we notice fractional numbers in the increase of a constituent, as for instance, 12 of oxygen in a second compound, while the first contains 8, which is the equivalent, as it is termed, by which chemists generally designate oxygen. The relation of 12 to 8, is as  $1\frac{1}{2}$  to 1, and of course involves a fraction. It is thought, however, and I incline to the opinion, that in proportion as analysis shall attain to perfection, all fractional numbers will be abolished, and thus the law of definite proportions, be more completely established.

Some persons have doubted, and continue to be sceptical, in

relation to this law of definite proportions. But the more thoroughly it is examined, the more correct will it appear to be; and while in its details, it is calculated to give to our science the certainty of mathematical demonstration, it offers some of the best proofs the world has ever beheld, in favor of the doctrine of an all-wise and overruling Providence, to the total overthrow and confusion of those who have vainly fancied, that this great globe is the offspring of chance. Not more certainly do the sun and the moon and the stars proclaim an intelligent, presiding Deity, than does the universality of the law of definite proportions, which reigns over the empire of matter.

Our attention will be more particularly directed to this interesting topic, in another portion of this volume.

The agents that have an influence on chemical action, sometimes operate in opposite directions; at one time they assist, while at another, they retard or hinder, combination. This diversity of action demands, on the part of the chemist, the most unremitting attention, for the neglect of which, he will be subjected to frequent disappointments.

We have seen that the force of cohesion stands opposed to affinity, and that it must be destroyed, in order to effect chemical union. But there are cases, in which the cohesive power retards combinations actually begun, and even puts a stop to them. It was this circumstance, that led the ancients into so many and so great errors. Even Bergman was not sufficiently aware of the energy of the attraction of cohesion, after the attraction of composition had commenced its operations; indeed, he viewed it as a dead force, or a power extinct. Berthollet, with more discernment, perceived that the force of cohesion was not annihilated, and that it sometimes resumed its activity, after the power, which had temporarily suppressed it, ceased to operate.

When we dissolve common salt or sugar in water, we effect a feeble combination; if the quantity of salt be sufficiently large, the water will be saturated, and there will be a perfect union between the two bodies. To this end, it is necessary that the water insinuate itself between the particles or atoms of the salt, in such a manner, as to overcome the force of cohesion which bound them together. Here, the solid is lost in the liquidity of the water. But if the force which has united the salt and water, be now diminished, by any circumstance whatever, such as the abstraction of part of the fluid, the solution will cease to be perfect. Just at the instant of saturation, affinity and cohesion exactly balanced each other; to destroy this state of equilibrium, it is only necessary to expose the solution to the air, for a given

period; or, if greater promptitude be desirable, we may place the mixture in a suitable vessel, and resort to distillation. Thus we soon effect the evaporation of a part of the liquid. The substance held in solution, by its affinity with the liquid, will thus be separated from the latter, or the latter from it, and its particles will again yield to the force of cohesion, which force had ceased to be operative, during the state of combination.

We sometimes see, that a body whose particles are reunited, after having been dissolved, is more solid than it was prior to solution. This occurs in the article of sugar, the particles of which do not cohere very firmly in the ordinary state. If the force which holds it in a watery solution, be destroyed by a slow evaporation, the particles will be reunited, and very hard crystals will be formed, which are known by the name of *sugar candy, rock candy, &c.*

While in some cases, we are forced to destroy the cohesive power, in order to effect combination, in others, we find that this very power destroys affinity; so that it may be affirmed of the attraction of cohesion, that it both favors and opposes combination.

The same is true of caloric, as we shall see more fully in the sequel. It exerts, on some bodies, an agency that tends powerfully to destroy cohesion, and it is often employed with this object in view. If we take a quantity of sulphur and lead, the cohesion of their particles will prevent the two bodies from uniting, although placed in actual contact; but if we destroy this attracting force, by means of heat, both substances will become liquid, and union quickly follows. But caloric not only urges forward the process of chemical combination; in many instances, we find it capable of destroying combinations already made. We have noticed, in a former remark, that mercury may be combined with a light and gaseous substance, as oxygen, and thus give rise to one of those compounds, usually called oxydes. To this end, we pour a given quantity of mercury into a matrass, having but a small aperture for the entrance of air; the stem of the vessel being too long to permit the escape of any portion of metal. This apparatus is exposed to the influence of a moderate heat, for the space of twelve or fifteen days. The changes that ensue, are too remarkable to admit of explanation, on any other than chemical principles. The mercury, from being nearly white, has passed into a red colored substance, and that, too, by combination with a colorless gas. The compound is designated by the term *precipitate per se*, to distinguish it from the *red precipitate* of the shops, which is formed from mercury and nitric acid.



In this process, it is plain that the heat employed has had a material influence. If the vessel, with its contents, had been exposed to the air only, and the agency of heat had been dispensed with, the lapse of months and years would not in all probability, have brought about a change. But the heat acting on the metal, separated its particles, destroyed their cohesion, and thus presented innumerable points of contact and so greatly favored the combination with one of the constituents of the surrounding air. But, if we take this same *precipitate per se*, place it in a retort, and subject it to the action of an intense heat, we shall find that the same agent, by which this compound was formed, will destroy the combination, and reproduce the metallic mercury, and separate the gaseous substance with which it was previously identified. Thus, under different circumstances, the same force, varying in intensity, promotes or destroys chemical union. To these opposite results, chemists have given the names of *composition* and *decomposition*, or *synthesis* and *analysis*. Exp.

Chemical combination is also influenced by the mass or quantity of the bodies subjected to its influence. It occurs, for example, that a substance will be more powerfully attracted by three, than by two parts of some other substance. Oil of vitriol is a compound of two simple substances, to which a given portion of water is always superadded. In the ordinary oil of vitriol, there is about one-fifth part of water. If we add three parts of water, beside that which is usually present, we make the acid and water to be in equal proportions. Expose the mixture to the action of heat, so as to effect distillation, and it will prove an easy task, to separate the three parts of water, which were last added; but to expel that portion which was originally combined, we must apply a much stronger heat. This latter portion of water was, therefore, united to the oil of vitriol, by a much more energetic affinity, than that which retained the other three parts; in other words, one molecule or particle of water is more firmly attracted by four particles of oil of vitriol, than three or four parts of water. Here, it is obvious, that the mass or quantity of water, has greatly influenced the strength of affinity.

Chemical combination or affinity is also affected by electricity and galvanism.\* Thus, various gases are made to combine, so as to form new compounds, simply by passing electric sparks, or a current of the galvanic fluid, through them, in proper vessels.

\* These are not mentioned here, on the supposition that they are distinct agents; my view of them, is accordant with the modern theory.

So also, these agents, acting upon water, decompose it into the gases of which it is formed.

Pressure is also influential in the operations of chemical affinity. A liquid and a gas, between which there is some strength of affinity, will combine, though only in certain proportions, under ordinary circumstances. In order to make the liquid take up or combine with more of the gas, the points of contact, already referred to, must be augmented, by continued agitation, and this is readily effected, by increased pressure. It is on this principle, that the artificial carbonated waters are prepared. This is one effect of pressure, but there is another. It may operate so as not only to retard, but prevent decomposition, as the following fact will demonstrate. Carbonate of lime was placed in a crucible, pressed down as closely as possible, and the vessel was made, as nearly as it could be, hermetically tight. Thus it was exposed to an intense heat, and the carbonate was not only undecomposed, but actually melted. If pressure had not been employed, pure lime would have been found in the crucible, the whole of the carbonic acid having been expelled by the heat, as is the case, in the ordinary mode of burning lime.\*

This fact also serves to explain an important geological phenomenon. It was formerly a source of wonder, that carbonate of lime should be found at a great depth in volcanic regions; it was supposed, that the volcanic heat must necessarily decompose all carbonates, and it would seem, that nothing but the counteracting power, of great pressure, could produce such a result. We proceed now to show the application of the forces or powers, already brought into view, to simple and double decompositions; or in other words, to exhibit practical applications of the laws of chemical affinity.

The most simple decomposition or analysis consists in the separation of two bodies, combined in one. As we have already observed, a compound is irreducible to its elementary parts, by mechanical means, and recourse must be had to chemical action. According to Bergman, a celebrated writer on affinity, the separation of two bodies, as they exist in a binary compound, (a compound of two) is effected by the addition of a third, only because the affinity of the latter, for one of the two bodies in the binary compound, is greater than that which held the two together. Hence, one of the two bodies is removed, and its place is supplied by the third; but as this is an important point, it is proper to illustrate it by experiment.

We suppose, therefore, that the compound  $A\ B$ , is formed

\* See Langier's *Cours de Chimie*.

by the union of two distinct bodies  $A$  and  $B$ . There is, of course, a certain degree of chemical affinity, subsisting between these two bodies, or they would not combine. But if we add another body,  $C$ , to this compound, we shall find that  $B$  is displaced, that  $C$  has united to  $A$ , forming the new compound  $A C$ . The first condition necessary, in order to be successful is, to have the force of cohesion diminished, so that affinity may exert its influence; and to make the operation as simple as possible, we shall suppose that the bodies are in the fluid state. The compound  $A B$ , being in a proper vessel, the solution of  $C$  is gradually poured in. The whole becomes turbid, indicating a *precipitation*, as chemists say, and showing that one of the constituent parts of the compound  $A B$  has been detached.  $C$  has actually joined  $A$ , the original compound no longer exists, but a new one, viz.  $A C$ , has appeared, and  $B$  having thus been liberated, is instantly subjected to the attraction of cohesion, and falls to the bottom of the vessel. If we desire to know the relative proportion of  $B$  to the whole compound, of which it was lately a part, we resort to a very simple process. We first shake the mixture and throw it on a filter. The new compound  $A C$ , being in solution, readily passes through the filter, and is collected in another vessel, while  $B$  accumulates on the filter, in the state of powder. This is, in the next place, to be dried, and sometimes a red heat is necessary, to free it entirely of moisture. If it be now weighed carefully, we shall know what proportion it bore to the original compound.

Having thus ascertained the weight of  $B$ , it might be inferred that the difference between  $A B$  and  $B$ , will denote the weight of  $A$ ; but chemists do not ordinarily proceed in this way. They endeavor to find the weight of  $A$ , by a course of procedure, similar to that pursued in the case of  $B$ . Thus, to the compound  $A C$ , dissolved in water, they add another, viz.  $D$ , which has a stronger affinity for  $C$ , than  $C$  has for  $A$ . A new compound  $C D$  is thus formed, and  $A$  is precipitated. By filtering as before, we collect  $A$ , and having dried it, we ascertain its weight. The affinity of  $C$  for  $A$ , in the first operation, by virtue of which the union of  $A B$  was severed, and the compound  $A C$  formed, is called *simple elective affinity* or *attraction*; because of an imaginary preference on the part of  $A$  for  $C$ , rather than for  $B$ . The same remark applies to the second operation, wherein  $D$  takes  $C$  from its combination with  $A$ , to form  $A D$ . Now, in the original compound  $A B$ , affinity was certainly operative, but the addition of  $C$  gave occasion to the action of a more powerful affinity, and hence we affirm, that  $C$  was capable of decomposing  $A B$ . Such, however, is not the play of affini-

ties under all known circumstances; for if  $A$   $C$  and  $B$  be taken in the solid form, and exposed to heat in a crucible, we shall find the original compound  $A$   $B$ , reproduced, while  $C$  will be detached or evaporated.

It would seem, that, in these cases, the result was not dependant on affinity, solely, but that other powers have been operative. In the decomposition of  $A$   $B$  by  $C$ , without heat, it is obvious that affinity owes nothing to the influence of caloric, but we cannot say that the force of cohesion has not performed its part. While  $B$  existed in chemical combination, it had no dependance on cohesion, but so soon as it was separated from  $A$ , its particles tended to reunion, and the force of cohesion thus operating, actually favored the combination of  $A$  and  $C$ , though this alone could not effect their union.

On the other hand, when we collect  $A$   $C$  and  $B$  in a crucible, and expose it to the heat of a furnace, affinity is influenced directly by temperature. By supposition,  $C$  is a volatile substance, whilst  $A$  and  $B$  are fixed bodies; of course, the effect of the heat will be to dissipate the former, and this very circumstance leads to the operation of the natural affinity subsisting between  $A$  and  $B$ , and they consequently recombine.

However tedious these details may appear, in the estimation of some readers, they are highly important, because they go to explain a fundamental truth, viz. that in estimating chemical union and decomposition, we must have respect to all the modifying circumstances, attendant on each given case. They, moreover, serve to confirm a remark formerly made, that the ancient definitions of chemistry, were imperfect. This science has for its object, not simply a knowledge of the action of matter on matter, but it takes cognisance, even to minutiae, of all the circumstances that affect the operations of affinity.

Defin.  
of  
Chem.

It will often occur in the simple decomposition of a body, as  $A$   $B$  by another,  $C$ , that  $B$  will not be precipitated. If this latter, although detached from  $A$ , does not yield to the attraction of cohesion, it will remain suspended or dissolved in the same fluid which contains the new compound,  $A$   $C$ . Here, decomposition is not very obvious, since the flocculi or floating particles, which ordinarily denote precipitation, are not seen in the liquid mass; but if we evaporate the solution, we shall succeed in separating  $B$  completely.

Another condition of  $B$  may present itself, which is important in the history of affinity. Thus, it may neither be precipitated nor remain in solution, but from its very nature, it may yield to the power of caloric and fly off in the state of gas.



Here, although the decomposition may be invisible, the peculiar odor of the volatilised matter may apprise us of the result. This is, of course, a third variety of simple decomposition, and is by no means, an unfrequent case.

*Double elective attraction*, is a term that has been applied to the phenomena, resulting from the admixture of two binary compounds, as  $AB$  and  $CD$ . If these be dissolved and poured into one vessel, there will be a double decomposition, and two new compounds will be formed, viz.  $AC$  and  $BD$ . But, in conformity to a position already defended, if we place these new compounds in other circumstances and operate upon them by means of heat, decomposition will again ensue, and at the close of the process, it will be seen, that the original compounds  $AB$  and  $CD$ , have been reproduced. We are not at liberty to say, that this is *elective affinity*. Let this case be illustrated by an example. The muriate of ammonia, (sal-ammoniac of the shops,) is the compound  $AB$ , and the carbonate of lime, (chalk,) the compound  $CD$ . If these be reduced to powder, intimately mixed, and then exposed to heat in a retort or other distillatory vessel, a mutual decomposition will follow; muriate of lime will remain in the retort, and carbonate of ammonia will be driven off, in form of gas. These new compounds, correspond to  $AC$  and  $BD$ . By proper means, the gaseous carbonate of ammonia may be condensed and procured in a solid state, for further experiment. Then take  $AC$ , the muriate of lime, and  $BD$ , the carbonate of ammonia, dissolve them in water, separately, and then mix the two. The result will be, another double decomposition, giving rise to the original compounds  $AB$  and  $CD$ . Here, it is obvious, that the action between these bodies, has been not only different, but entirely opposite; and this variety has been occasioned by diversified action of heat, the abstraction of heat and solution. Hence, it would be improper to attribute these results, altogether, to affinity, because that being the only agent, the products should be uniformly the same, which is contrary to experiment.

There is another variety of affinity, which has some importance attached to it, and is usually called *complex*. This title has been given, because the results are somewhat paradoxical, although full of instruction. Thus, if  $C$  be added to  $AB$ , the latter remains unchanged; if to another portion of  $AB$ , we add  $D$ , no effect is perceptible. Let  $C$  and  $D$  be joined in the compound  $CD$ , and add this to  $AB$ , and a double decomposition takes place. To illustrate this problem, I take nitric acid and pour it into a solution of sulphate of soda, (*Glauber's salts*,) but the latter still retains its transparency. To another portion of

this saline solution, I add lime-water, and yet there is no alteration. Then I mix nitric acid and lime-water, to form nitrate of lime, and on adding this to the solution of sulphate of soda, both are decomposed; sulphate of lime and nitrate of soda are formed.

Again; there is a species of affinity, to which the term *pre-disposing* or *disposing* has been given. The results of this affinity, differ from those of double decomposition. In the latter, two compounds act mutually on each other, and two *binary* compounds, unlike the original, are produced. In predisposing affinity, *three* of the bodies unite and give rise to *ternary* compounds, while the fourth escapes in the state of gas.

If we take any quantity of iron filings and add a portion of water to them, and even call in the aid of heat, we shall not discover any action between the metal and the fluid; but if we slowly pour in oil of vitriol, a violent action ensues. The water is decomposed, its oxygen unites to the iron, forming an oxyde of iron, which combines with the sulphuric acid (oil of vitriol,) and forms the sulphate of iron, (green copperas of the shops). Here, the sulphuric acid, iron, and oxygen of the decomposed water, are united in a ternary compound, (a compound of three,) while the fourth body, viz. hydrogen, the other component part of water, flies off in form of gas.

Sometimes a ternary compound is formed, when one body is added to two others, without the evolution of a fourth. Thus, oil and water will not combine, although agitated for a great length of time. If a strong alkaline solution be poured in, as for example, the well known spirits of hartshorn, a gentle agitation forms the whole into an uniform mass. Here the alkali is called the *disposing* agent, or that which occasions the union.

Exp.

## CHAPTER II.

Of Imponderable Bodies. Light. Caloric. Electricity. Galvanism. Magnetism.

CHEMISTS have usually divided their subject into two great classes, viz. *imponderable* and *ponderable*. The former embraces but few objects, all of which are destitute of weight, and are chiefly understood by their effects. The latter class includes all the bodies in nature, on which chemical affinity can exert an influence, and is, of consequence, very extensive in its range.

Of the imponderables, there is one which has but little interest for the elementary student of chemistry, and that is, *magnetism*. It seems to place itself most fitly in the province of natural philosophy, and there we shall permit it to remain.

To electricity and galvanism, which are now regarded as convertible terms, a distinct chapter is appropriated, in part fourth of this volume.

*Light* exerts a considerable degree of action on chemical combinations; but of all the imponderables, caloric is the one which claims our chief attention, and it shall be particularly noticed. At present, however, a few remarks seem to be called for, on the relation of light to the phenomena of chemistry.

The indispensable necessity of light, to perfect vision, is well understood. Newton supposed it to be an emanation from the stars, while Huyghens thought it was a fluid, diffused through immeasurable space, whose vibrations are propagated by a succession of undulations. This opinion has acquired not a few advocates, in modern times; but whether it be true or false, is of little importance to the chemist as such.

Light cannot be concentrated into a perceptible mass, by any force that we can employ. Its levity is equal to its rapidity; and the one is probably the consequence of the other. Although it passes through space, at the immense rate of 200,000 miles in a second, it does not, in the smallest degree, injure the healthy eye, although formed with so much delicacy.

Light acts on bodies in several ways. Sometimes it falls back on our vision, after having been arrested, at a given surface; it is then said to be *reflected*; sometimes, passing through bodies, its course is turned a little from the straight line, and then it is said to be *refracted*. The luminous rays may be partly reflected and partly absorbed. If all the rays are reflected, we have the sensation of *white*; if, on the contrary, all the rays are absorbed, we have the sensation of *black*; and between these two extremes, we find the various colors. A body is red or blue, just as it reflects the red or blue rays, and absorbs all the others.

The term *transparent*, is applied to all bodies, through which the light readily passes; this fluid is decomposed, that is to say, its rays are separated, when it is made to pass through a transparent piece of glass, called a *prism*. The decomposition, thus effected, presents the following order of the rays, which forms a word that is somewhat peculiar, but which, if recollected, will at once remind the student, of the order in which the rays of light are separated. Thus we have, violet, indigo, blue, green, yellow, orange, red, which, by simple juncture, make the word, *VIBGYOR*.

Light is necessary to the healthy functions of all beings. Men who are deprived of it, for a considerable space of time, such, for instance, as the inhabitants of subterranean places, are usually feeble and sickly. Even the vegetable kingdom is

deteriorated by the defection of light. If a plant be deposited in a deep cave, it will invariably incline to that part of the cave, which is nearest to the external light, thus evincing an instinctive fondness for this fluid.

Vegetables require the action of light to enable them to evolve a gaseous matter, which, in excess, would be hurtful to them. They absorb from the surrounding air, a large portion of carbonic acid, which they decompose, retaining the carbon of that acid, for their own use. But in order to gain this carbon, which is essential to vegetable life, it is requisite, that the oxygen, with which it is combined in the carbonic acid, should be separated. This result is accomplished by the agency of light, which enables the vegetable to decompose the acid and set the oxygen free. Thus vegetation performs a most friendly office in behalf of animal life. It takes up a recrementitious matter, exhaled from the lungs, and which is unfit for respiration, and after resolving that matter into its elementary parts, sends out a perpetual stream of vital gas, (oxygen,) to preserve, in purity, the atmospheric air.

Chemistry appreciates the importance of light, in a variety of respects. Several of the metallic oxydes, (compounds of oxygen and metal,) are decomposed by the action of light, as for instance, the oxydes of silver and gold. The lunar caustic of the shops, when first formed, is almost white, and may be kept in that state, if light be excluded, but it soon changes color if this precaution be neglected. Phosphorus is, in like manner, tinged of a deep yellow, by exposure to light.

If aqueous chlorine be exposed to a strong light, decomposition of the water ensues, and muriatic acid is formed. Nitric acid, of the strongest and clearest sort, is changed from the same cause, and it becomes yellow, in a few days, even in well stopped bottles.

Some of the effects, attributed to the agency of light, are, by many persons, ascribed to caloric, and the question is still at issue, whether these fluids are not identical. It is abundantly plain, that caloric, as well as light, is reflected by polished surfaces, and that in both cases, the angle of reflection is equal to the angle of incidence. A body strongly heated, becomes red, incandescent, luminous. When light is concentrated by a burning glass, it gives rise to all the effects connected with heat. Resemblances of this sort, have inspired the belief, that light and heat are modifications of one and the same substance.

Some experiments go to establish, between light and caloric, this identity of nature and action. If we mix chlorine in the gaseous state, with an equal volume of hydrogen, and expose the



mixture to the contact of a vivid light, sudden combination will ensue, and frequently with a degree of energy that subjects the operator to some hazard. But caloric alone, independently of light, is adequate to the same result. Into a vessel containing the same mixed gases, plunge a piece of hot brick or any other heated body, and explosion will follow. So also, if you write on two pieces of linen cloth with the indelible ink of the shops, and expose the one to the light of the sun on a cold day in winter, and the other to the heat of a close stove in a dark room, precisely similar results are obtained. The writing, in dark colored letters, will be perfectly legible. Here it would seem that caloric and light act precisely alike, and thus far they are identified.

*Of Caloric.* All bodies in nature would assume the solid form, if the fluid or substance, called caloric, did not perpetually operate, by its repelling power, to overcome the force of cohesion, and so give rise to liquids and gases. The particles of matter, yielding to the inherent power with which nature has endowed them, tend at every moment to solidity; but caloric interposes, and its repulsive energy drives them asunder. It is owing, in no small degree to the balancing of these opposing influences, that the harmony of the universe is maintained.

But what is caloric? It indicates by its derivation, from *calor*, nothing more nor less than heat. Yet philosophers have preferred to intimate by the term caloric, a cause of which the sensation called heat, is the proper effect. On our approach to a fire, we experience a sensation that is familiar to every child, but ought we to confound that sensation with its cause? To this query, philosophers very generally give a negative reply. Some have alleged, that heat resulted, simply from the intestine motions of the particles of matter; others conjecture, that caloric is a real substance, which, some how or other, gives rise to the sensation of heat. To the chemist, it is not essentially requisite to determine which hypothesis has the nearest connexion with truth. His business, more particularly relates to the obvious effects of caloric, than to its intrinsic nature.

It is agreed, universally, that this important agent is extremely subtile, incompressible, imponderable, and elastic. It is further conceded, that it exists in two different forms, viz. *free* and *combined*. We affirm that caloric is in the free state, when in its passage from a stove or fire place, it occasions a feeling of warmth. Escaping from its source, as from a focus, it sheds its influence on all surrounding objects, and we are perfectly conscious, that such is the fact. It moves in straight lines, precisely as light does, is reflected by polished surfaces, and has received

the appellation of *radiant caloric*. When caloric passes to bodies from a stove, they are warmed by *radiation*, and as the nature of their surfaces prevents the reflection of those rays to some other objects, they are said to be absorbed or to combine with those bodies. If, however, the rays of heat fall on a polished metallic surface, they may even travel through its substance, without increasing the temperature; for, to polished metals belongs, peculiarly, the power of reflection. Hence, a bright tin bason may be held in front of an open fire, without being heated, while the hand of an individual, placed equally near to the focus of caloric, will be almost burned. The tin bason reflects the rays of heat back to their source. Hence, too, highly polished brass andirons will remain exposed to a blazing fire, with very inconsiderable augmentation of temperature, while a beef steak, suspended between them, would be roasted. We shall have occasion to notice these phenomena again.

There is a peculiar feature of caloric, to which some writers have given the name of *tension, elasticity, expansibility, &c.* Its power of penetrating bodies in every direction, of radiating from their surfaces, its tendency to escape with rapidity, under certain circumstances, all these properties give it the character of expansiveness, or tension, and this is augmented as temperature increases. This peculiarity seems to be dormant, ready to be excited into action by an appropriate stimulus; such, for instance, as a sudden change in the temperature of bodies.

When substances are heated, they are in a condition to yield caloric to surrounding objects; if they are at a reduced temperature, they seem to be peculiarly fitted for the absorption or reception of caloric. Now this twofold effect is justly attributable to another property, which is denominated, the *tendency to equilibrium*. There is, therefore, a power of diffusion or transmission, inherent in caloric, whereby heated bodies part with a portion of that fluid; and there is also a faculty of concentration or absorption, whereby substances of inferior temperature, acquire an increase of heat.

By way of illustration, let us take two metallic balls, that are placed on a mantle, directly over an open fire. They are, of course, at the same temperature, with the air of the room; and so long as both balls continue in this condition, each will have its proper quantity of caloric. In other words, a state of perfect equilibrium exists, and will continue, until some adequate cause shall disturb it. Let one of the balls be placed in the fire, and there acquire an addition of from one to two hundred degrees of temperature, and then place it on its stand, in the vicinity of the other ball. Now, there is a lack of that

equilibrium, which was so apparent, a short time since, and a struggle is commenced and carried on, to restore the recent condition of things. The hotter ball gives out rays of caloric to all surrounding objects, and the other ball is constantly receiving its share; and this process of giving and receiving, does not cease, until a perfect equilibrium is regained. Reverse the case, and similar phenomena present themselves. Suppose that one of the balls, in place of being greatly heated, is surrounded by a freezing mixture, until its temperature is brought down to zero. In this condition it is placed near to the other ball, which being, relatively, the hotter of the two, sends out rays of heat, of which the colder ball is constantly receiving a due proportion, and this phenomenon continues, until both balls exhibit precisely the same degree of temperature. In this way, all bodies are gaining caloric, just in proportion as others are losing it, and thus nature presents a perpetual strife in matter, the aim of which appears to be, the establishment of a perfect equilibrium.

In addition to the faculty of *giving* out and *taking* in, to which we have just referred, there is another, to which these are subordinate, viz. the *reflecting power*. Allusion has already been made to this faculty, but it is proper to notice it more particularly. We have seen that metallic bodies, when unpolished, absorb heat and evince an increase of temperature; and that when colder bodies, in their vicinity, become heated, it is by virtue of caloric radiated from the hotter body. Highly polished metallic surfaces, upon which rays of heat fall, instead of being warmed thereby, reflect the caloric, thus sent to them, upon other bodies.

To illustrate the *reflecting power*, recourse has been had, to a very simple contrivance. Two concave mirrors, made of highly polished tin or sheet brass, are placed on suitable stands, directly opposite each other, at the distance of from five to twenty feet. In the focus of either mirror, an eight ounce or pound iron ball, heated nearly to redness, is fixed on a suitable stand; in the focus of the other mirror, we place a common thermometer, or what is preferable, a differential thermometer. No sooner is the latter in its proper focal spot, than the action of the reflected heat is seen in the movement of the thermometrical fluid. Exp, If in place of the thermometer, we suspend an ether glass, half full of sulphuric ether, expansion ensues, and the fluid is driven out at the capillary point, and if a lighted Exp. taper be applied, a beautiful experiment is made. The vaporised ether is instantly fired, and a brilliant arch of flame continues, until all the ether has passed out.

In lieu of the thermometer and ether glass, a candle may be



employed, having its wick exactly in the focus of the mirror. The threads of the wick must be separated, so as to make room for a piece of dry phosphorus, as large as a pea. Things being thus adjusted, the heated ball is fixed on its stand, at the opposite mirror, and, almost immediately, the phosphorus is inflamed and the candle is lighted. (See Plate 2d, Fig. 4.)

But, it is demanded, what reason is there for supposing, that these results do not follow from the simple radiation of heat and its well known tendency to equilibrium? A little change in the experiment, will furnish a sufficient answer to this question. Let the mirrors remain in their place, and also the thermometer or ether glass or candle, (as it may be,) in its place. Then, instead of fixing the heated ball in the focus of the opposite mirror, suspend it on the focal line drawn from mirror to mirror, but equidistant from both. Although the ball is now much nearer to the thermometer or ether glass, than in the first experiment, it acts so feebly, if at all, that no change is produced. Yet, if the results of the other experiments depended on radiation or tendency to equilibrium, alone, there should not be a failure in the present instance.

The focal ball does, unquestionably, radiate caloric; for, as it is a dull, unpolished body, it cannot reflect. The best reflectors viz. polished metals, are the worst radiators; and the best radiators, as all dull, tarnished bodies, are the worst reflectors, or rather, they do not reflect at all. But what becomes of the caloric radiated or sent out, from the focal ball? We answer, it passes or diverges to the circumference of the corresponding mirror, from which it flies off in straight and parallel lines to the circumference of the opposite mirror, from which, by a natural tendency, it converges to the central or focal point, and there falls on the thermometer, ether glass, or candle, producing the effects previously stated.

The experiment with the concave mirrors, may be further varied, to illustrate the reflection of heat. Thus, substitute a lump of ice, for the heated iron ball, and very soon, the thermometer in the focus of the other mirror will be affected, and its fluid will fall. Now, say some, this is proof, that cold is a positive agent, and not the mere abstraction of heat. They contend, that the frigorific particles pass from the ice to the thermometer, and act on it, by a positive power. This, however, is a mistake. Remove the ice from the focal point, to a place midway between the mirrors, and the thermometer will not be affected; whereas, on the supposition of a positive frigorific agency, the reverse should occur. The true state of the case is this. The ice, being in its proper focus, receives

Exp.

caloric from the thermometer, which is comparatively, the hotter body. The rays, passing from the thermometer to the circumference of its mirror, then go in straight parallel lines, to the circumference of the opposite mirror, and thence converge, to meet in the ice, which occupies the focal point. The fluid of the thermometer sinks, therefore, because it parts with its caloric, to give it to the colder body.

This *reflecting power* of caloric is susceptible of useful applications. Hence, the practice of interposing between the frame work of churches and stoves, large pieces of polished tin. If black and rough sheet iron were used, in the room of the polished metal, the evil would not be lessened, but increased; for a good radiator of caloric, would thus be placed nearer to the wood, than the stove itself. The tin acts favorably, by reflecting the rays of heat back upon the stove, and so screens the wood from the influence of an intense heat.

We notice another highly important property of caloric, viz. its effect in changing the condition of bodies, as from solid to fluid, and thence to gas. When this result is not fully accomplished, because of an insufficient degree of heat, we realise an augmentation in the bulk of bodies, a *dilatation* or *expansion*. Many instruments have been constructed, to evince this property. Take a brass ball, turned so as to be perfectly spherical, and have a handle accurately fitted to it; and procure an iron ring, the aperture of which shall be just large enough, to allow the ball to pass, when cold. Dip the ball into boiling Exp. water and then apply it to the ring. The caloric of the hot water, has so expanded the brass ball, that it will not enter the ring, much less pass through it. An iron bolt and ring are also used for the same purpose. The former, which can easily be passed through the ring, when cold, is so enlarged by the action of a red heat, that it cannot be made to enter.

The wheelwright practices, daily, on this principle. In making a tire or hoop for a wheel, he does not calculate to have it large enough to embrace the wheel, when cold. He makes it of a diameter considerably smaller than that of the wheel; and the reason is, that when the tire is made red hot, its diameter is increased, by the reason of the entire expansion of the metal. In this state, he places it on the wheel without difficulty, and then taking advantage of the opposite of the law of expansion, viz. contraction, by the abstraction of heat, he dashes cold water on the tire, and thus forces the particles of iron back to their original state of contact.

The side walls of an immense edifice, in France, were forced out from the perpendicular line, by the pressure of the roof,

and fears were entertained, that the building would be destroyed. An ingenious application of the expansive power of caloric, and its opposite, viz. contraction, saved the edifice. Holes were drilled a few feet below the edge of the roof, and at proper distances from each other, in both walls. Thick bars were then passed through these holes, across the building, having the ends to project several inches beyond the walls, and each end wormed into the form of a screw, with a large nut or burr attached, and this screwed up to the wall. Then directly under each bar, suitable fixtures were placed, to make the iron red hot, and of course to expand it. During this operation, screw wrenches were applied to all the nuts, with as much force as could be exerted. After this, contraction was effected by the action of cold water, the wrenches being all the while powerfully employed. Every act of heating and cooling, conjoined with the use of the wrenches, produced a visible effect; and by perseverance, the walls were brought up to the straight line; and the historian tells us, that by continued efforts, they might have been forced within the building.

On the law of expansion, the formation of all the instruments usually called *thermometers*, is dependant, whether air, alcohol, or mercury be used in their construction.

A *thermometer* means, literally, a measurer of heat, and as continual reference is made to this instrument, in all our books of science, it is proper to give the subject a little consideration.

Sanctorio has been regarded as the inventor of the first thermometer, of which we have any knowledge. It is called, very frequently, the air thermometer, because the external heat acts on the air contained in the bulb of the instrument, and by expanding it, causes the colored fluid to sink. One of these instruments may be made very easily, by taking a thermometer tube of large size, embracing the bulb with the palm of the hand, and then immersing the stem in a vessel containing black or red ink. Let the hand be immediately removed and the ink will rise in the stem, nearly half way to the bulb. Thus prepared, the instrument wants only a scale of degrees to complete it. It is plain, that as the air contained in the bulb, expands just in proportion to the external heat, and in the same proportion, causes the colored fluid to fall, the thermometer of Sanctorio will serve to show the changes of temperature tolerably well. An important objection to this instrument is, that it is affected, materially, by the pressure of the atmosphere, and on this account, chiefly, it was soon laid aside.

The thermometers now in use, are uninfluenced by atmospheric pressure, and they indicate an increase of temperature, by the

expansion of their contained fluids, viz. either alcohol or mercury.

The principal reason for a preference that was once given to alcohol, was, that it bore the lowest temperature, without freezing, and herein, it served a better purpose than mercury. On the other hand, the more gradual expansion of mercury, rendered it greatly superior to alcohol, for measuring very elevated temperatures. In addition to this, mercury seems to expand more accurately, in proportion to the increments of heat, than any other fluid.

There are several varieties of the mercurial thermometer, among which are those of *Fahrenheit* and *Reaumur* and the *Centigrade* thermometer. And as a table is given at the end of this volume, by which the differences between these instruments can be learned, as well as the modes of converting them into each other, it will be unnecessary to give a minute description of each, especially as they agree, in the principles of their construction. I shall, therefore, present an outline of the formation of a *Fahrenheit's* thermometer, only, as that is the instrument generally used in this country.

In forming this instrument, it is requisite to have a tube whose bore is of uniform size throughout, and that is, ordinarily, less than a line. The tube is blown into a bulb at one extremity, and the mercury is introduced, by heating the bulb over a spirit lamp, to expel moisture and air. The next thing to be done, is to dip the stem in mercury and to cool the bulb, so as to form a vacuum, whereupon the mercury rushes in with rapidity. Now, it is all-important, before proceeding to close the open end of the tube, which is called, *hermetically sealing* it, that the excess of mercury be expelled and all the air driven out. To this end, the spirit lamp is again applied to the bulb, and at the same time, the end of the stem is softened by another lamp, and the operator, seizing the proper moment, closes the aperture in the stem. It will now be found, that on inverting the tube, the mercury will fall from the bulb, down to the end of the stem, thus evincing that the air has been effectually excluded, and, of course, that the mercury will not be influenced by atmospheric pressure. The tube being thus prepared, the next step is to graduate it correctly. To do this, we have first to obtain two fixed points, which shall be the same in every Fahrenheit thermometer. In order to accomplish this end, Sir Isaac Newton's method is usually pursued. It is based on the fact, that when a thermometer is plunged into ice, in the act of dissolving, or into water that is boiling, the elevation of the mercury will be the same in all countries, under equal circumstances. To obtain the



freezing point, the instrument is immersed in snow or pounded ice, which is gradually liquifying, till the mercury becomes stationary. The boiling point is not so easily procured, because the ebullition of water may vary, according to circumstances. The water should be as pure as it can be had, and not over an inch in depth; it should be boiled briskly, in a deep metallic vessel, so that the stem of the instrument may be surrounded by an atmosphere of steam, while the bulb is in the water. The barometer should stand at thirty inches.

Having thus ascertained the freezing point,  $32^{\circ}$ , and the boiling point,  $212^{\circ}$ , the scale is to be divided into one hundred and eighty equal parts or degrees, between those two points. And as the zero of Fahrenheit is  $32^{\circ}$  below the freezing point, the scale should be divided into that number of equal parts, below  $32^{\circ}$ . It is also occasionally useful to have a thermometer tube adapted to a scale, so as to carry the degrees down to  $39^{\circ}$  below zero, at which point mercury freezes; and up to, or even above,  $600^{\circ}$ , at a little over which, mercury is changed to vapor.

The *hinge* thermometer differs from that just described, in but one respect. At about three inches from the lower end of the scale, there is a transverse cut, and the piece, thus detached, is fastened to the other portion, by means of two small hinges. This contrivance is resorted to, in experiments with acids, when it is necessary to dip the thermometer into the corrosive liquid. By turning back the hinged piece, no injury can result to the instrument.

There is no essential difference between the thermometer of Fahrenheit, and the other two, already named. In Reaumur and Centigrade, 0 indicates zero. The boiling point of the former is  $80^{\circ}$ , and of the latter  $100^{\circ}$ ; and the interspace between zero and the boiling point, is divided into eighty equal parts in the one case, and into one hundred in the other. The last species of thermometer to be noticed, is called the *differential* thermometer of Leslie, (Fig. 3d. Plate 2d). It is essentially an air thermometer, and differs from that of Sanctorio, mainly, in the exclusion of the agency of atmospheric pressure. Although frequently called *Leslie's* thermometer, it was known in the year 1676, more than a century before that individual was born. It consists of two very thin glass balls, blown out into a uniform shape, and connected by a tube, bent twice at right angles, so as to resemble the letter U. The greater part of the tube is filled with sulphuric acid, colored with carmine, and the balls contain common air. Now this instrument, from its very nature, cannot indicate the temperature of a room, in which it is placed, because both balls are acted upon alike. But the slightest differ-

ence in the temperature of the two balls, is instantly seen; for the air in the hotter ball, being expanded by the heat applied, drives the colored liquid up into the opposite ball. Hence, this instrument is admirably adapted to the experiments with the concave metallic mirrors, and this is, perhaps, the only use to which it is applied in chemical lectures.

To measure exceedingly intense degrees of heat, *pyrometers* have been constructed. Of these, the instrument of Wedgwood, is best known; but as its indications, like those of all pyrometers, are uncertain, it is seldom used, at present. It is founded on the property which clay possesses, of contracting, when strongly heated, without resuming its former bulk, as it cools. But as experiment has shown, that the alumine, (the base of the clay,) cannot be freed, entirely, from water, by any degree of heat, it is probable that a part of the contraction, at least, is owing to the gradual dissipation of the water or moisture of the clay. The apparatus consists of a metallic groove, twenty-four inches long, the sides of which diverge, from the width of three-tenths of an inch below, to a half inch above. The clay, made into forms to slide in the groove, is heated to redness. The subsequent contraction of the pieces of clay, is determined by allowing them to slide from the top of the groove downwards, till they arrive at a part, beyond which they cannot pass. Wedgwood's groove is divided into  $240^{\circ}$ , each of which is said to equal  $130^{\circ}$  Fahrenheit. His zero corresponds to the 1077th degree of F.

The most obvious effect of caloric, is to change solids into fluids, and these into gases; and hence, bodies have been arranged under these three divisions. Not only will the addition of heat alter a substance from the solid to the fluid state, and thence to the form of gas, but the abstraction of heat will reduce bodies from the gaseous to the liquid, and finally to the solid form. The facility with which bodies pass from one to the other of these states, varies very considerably, and depends on several circumstances, of which density is not the least important. Lead, tin, and bismuth, which are naturally solids, are changed to the liquid state, by a temperature, which is not very much elevated. Platina, also a solid, is bound so firmly by the attraction of cohesion, that it resists the most powerful furnaces. Caloric seems to separate the particles of metals, that are of easy fusion, as water detaches the particles of a salt from each other. As it effects this separation, it gives a degree of mobility to the particles, and thus fluidity is occasioned.

If we take another metal, which is found in the fluid form, viz. mercury, and expose it to a heat of  $650^{\circ}$  Fahrenheit, its particles are gradually separated and refined, until they assume the gaseous state.

At the moment in which highly heated bodies begin to experience a change of condition, the caloric received by them, no longer serving to increase their temperature, merely dilates them or increases their bulk, and this goes on, until the change is completely affected. No substance passes more readily through the several conditions of which we have been speaking, than water, and of course, it offers a fair sample of the effects and nature of these changes. At  $32^{\circ}$ , or a little below, we find it in the state of ice; to reduce it to the fluid state, heat must be added in sufficient quantity, equal to that which the water lost in becoming solid. If we take a pan of ice or snow at  $32^{\circ}$  and

place it over a fire, some time will elapse, before fusion takes place; and when that has been accomplished, a thermometer, immersed in the liquid, immediately after fusion has occurred, will stand at  $32^{\circ}$  although it is certain, that a large quantity of caloric must have been taken in by the ice, while changing its condition. So also, if a pound of water at  $172^{\circ}$ , be poured on a pound of ice at  $32^{\circ}$ , the ice will soon dissolve, but a thermometer placed in the liquid, will stand only at  $32^{\circ}$ . Of course,  $140^{\circ}$  of heat have disappeared, and there is no way of accounting for this apparent loss, but by supposing, that a given amount of heat must be absorbed or taken in by every solid, in order to pass into the fluid state. The caloric, therefore, that disappears on adding hot water to ice, is called the *caloric of fluidity*, because it is essential to the change spoken of; and, as this caloric is concealed, and cannot be detected by the thermometer, it is also called *insensible* and *latent heat*.

Suppose we subject water to the process of boiling; we elevate the heat, until the particles, being converted into vapor, begin to escape. So long as this change is going on, it would be idle to add more fuel, because the temperature of the water cannot be increased. Hence it is well known, that the boiling point of water is  $212^{\circ}$ , beyond which, it is impossible to heat that fluid. All superfluous heat, that is, all heat over  $212^{\circ}$ , is spent on the dilatation of the vapor or steam; and the expansion, thus effected, is truly enormous, for the vapor of water, thus dilated, may be made to acquire a bulk 1700 times greater than that of the water which it represents. Now, the excess of heat, which thus enlarges the original volume of water, does not effect the temperature, in the smallest degree, after the fluid has been raised to  $212^{\circ}$ ; where then is the extraneous caloric? I answer, it is hidden or become latent, in the expanded vapor.

These remarks apply, with equal force, to the fusion of solids, that are harder than ice. Thus, while lead and tin are passing into a state of fluidity, caloric is becoming latent, or, it is ab-



sorbed and scattered between the solid particles, so as to repel them from each other, and thus put them in that state of mobility, which must precede the change to liquidity. The caloric, thus acting, does not alter the temperature, because it spends its force, as a principle of repulsion, in bringing about a change of condition. This being accomplished, a further portion of caloric, soon indicates that appropriate effect which is denominated, *sensible* or *free caloric*. The heating power of caloric is, therefore, distinct from the faculty of altering the condition of bodies. When a metal begins to melt, the fused particles do not experience an increase of temperature, until every atom is melted, and, for an instant, the heat is stationary. When the whole is completely fluid, then it acquires, rapidly, the temperature of the surrounding fire, and may even be so far elevated, as to pass into the state of vapor. All heated bodies seem to have a disposition, to send out their caloric to other bodies; or rather, this tendency of escape, seems to be characteristic of caloric itself, and hence we have the term, *conducting power*, and this property is exceedingly unequal in different substances, and in some, scarcely exists at all. Thus, you may take a piece of charcoal in your hand, with impunity, the remote end of which is burning; and, in like manner, a stick of sealing-wax may be held, while a part of it is blazing. Now these and similar bodies are said to conduct heat badly, and this very defect has some useful applications. The metals, on the other hand, are excellent conductors. Glass is said to hold a medium rank, in the list of the conductors of heat.\* By a very simple contrivance, the difference of conducting power, in different bodies, may be happily illustrated. Take a piece of copper, a foot square, or, giving it a circular form, let it be a foot in diameter. Elevate it by props of iron wire, on a board of corresponding size, so as to allow room for a spirit lamp, to heat the lower surface of the copper. Then, at equal distances, say four inches apart, place a small piece of glass, an

Exp.

\* The imperfect manner in which glass conducts heat, is frequently a source of inconvenience and loss. Hence, glass utensils of considerable thickness are fractured, by the application of hot water, either within or without. A tumbler, placed in hot water, in cold weather, will be broken, because the heat expands the external layer of the particles of glass, while, on account of imperfect conduction, the internal particles are not, in the least degree, heated; the opposite conditions of expansion and contraction, within and without, cause the fracture. On the same principle, a glass vessel is broken, by pouring hot water into it, while the outer surface remains cold. For this reason, we employ the thinnest glass vessels that can be procured, for chemical purposes. The thinner the glass, the more readily is caloric conducted.

other of cork, a third of lead, a fourth of zinc, a fifth of porcelain, &c., &c., having the several pieces of equal thickness. Then, on each piece lay a portion of dry phosphorus, about half the size of a pea, and apply a lighted taper to the spirit lamp. It will be found, that the phosphorus will be kindled on the different pieces, precisely in agreement with their several powers of conducting caloric. The experiment is pleasing, as well as instructive.

Another contrivance is used, to illustrate the same faculty of caloric. Rods are procured, of glass, iron, copper, silver, wood, &c., of equal length and thickness, and passed through a board or plank, with each end projecting the same distance from the wood. On one end of the rods, a coating of beeswax is placed, and the other is immersed in boiling water. The wax will melt on the different rods, agreeably to their several conducting powers.

Liquids are bad conductors of caloric; indeed, it has been doubted, whether they, in any degree, merit the name of conductors. Some have supposed, that because water, placed over a fire, gradually becomes heated, and finally boils, it exhibits the *conducting* power of caloric. But there is another and better solution of the phenomena of ebullition. Heated particles, both of air and water, have a less degree of specific gravity, than cold particles. The fire acts on the lower or more depending part of the liquid, whereby the particles are heated, and as a consequence they rise to the top of the vessel; the cold particles descend, as the heated particles rise, and this descent of cold, and ascent of hot particles, continues until the whole mass is heated. In all this, there is no evidence of conducting power, and the result depends on the agency of caloric, in changing the specific gravity of fluid particles. The same process occurs in heating a large room by means of stoves. The air of the room is a gaseous substance, having fluid properties. Every particle of that air, when heated, is reduced in its specific gravity, and rises to the ceiling or roof, while the upper and colder particles fall; these being heated, in their turn, rise, to make room for other particles. Thus, the air of a large room is warmed, and the conducting power of caloric has no agency in the operation.

A very simple experiment will serve to prove, that caloric is not conducted by water downwards, however it may operate in the contrary direction. Take a florence flask, the bottom of which has been removed, or in lieu of this, a common  
 Exp. tin or glass funnel. Pass the stem of an air thermometer down the neck of the flask or funnel, and lute it in that po-

sition, having the bulb projecting upwards. Then pour in water, sufficient to cover the bulb of the thermometer, and on this, a small quantity of ether. If this inflammable liquid be now kindled, it will be seen that no effect is produced on the fluid in the stem of the thermometer, while active combustion is going on, in its immediate vicinity.

It seems to be a fixed law of nature, that bodies which conduct heat well, are not able to retain it, and that they diffuse it as readily as they receive it. Thus, if we desire to keep a substance quite warm, as for example, water, we should not place it in a metallic vessel, but much rather in glass or earthen vessels.

We have seen, already, that charcoal is a bad conductor of heat, and on this principle, coffeepots with double walls, (or, as it were, one coffeepot inside of another,) and finely powdered charcoal filling the interspace, have been found, by the lovers of hot coffee, to be admirably adapted to preserve that beverage, at a proper temperature, for several hours.

In addition to the attraction of cohesion which holds particles together, and the repulsion of caloric which tends to separate them, there is a third power, which modifies the elastic energy of caloric. In some substances, especially the more volatile, there is such a tendency to expansion, that they would always appear in the æriform state, if an external cause did not oppose this separation of their particles. This modifying power is the pressure of the atmosphere; and it operates also to prevent water, which is not classed among volatile matters, from changing to vapor at a temperature under  $212^{\circ}$ ; and the same external cause, renders a heat of more than  $600^{\circ}$  necessary, in order to convert mercury into the gaseous state.

To illustrate these positions, we have only to place a vessel of water, at  $100^{\circ}$ , under the receiver of an air pump, and then exhaust the latter, as much as practicable. The effect of this operation is, to remove from the surface of the water, Exp. the weight of the air, and this being accomplished, the water instantly boils.

An experiment, of a pleasing kind, was invented by Berzelius, calculated to confirm the foregoing proposition. He took a long necked matrass, and filled more than half of its ball, with hot water, and having closed its mouth, he in- Exp.verted it. Vapor or steam soon filled all the space not occupied by the water; and although the weight of the air could not have an influence, yet the water did not boil. The reason was obvious; the vapor itself performed the part of the atmospheric air, and by its pressure on the water, opposed its ebulli-

tion. To prove that the vapor really exerted this influence, he applied a sponge containing cold water, over the spherical bottom of the mattrass, so as to condense the vapor, and thus cause a partial vacuum. In an instant, ebullition was excited.

The common *pulse* or *palm* glass, also shows the effect which an abstraction of atmospheric pressure has on boiling. It is constructed with two bulbs, joined by one common stem; one of the bulbs contains a portion of colored ether or alcohol, and the aperture is closed, so as to exclude the air entirely. On applying the palm of the hand to the bulb, containing the colored liquor, the fluid is driven over to the opposite bulb, and before the whole has escaped, a violent motion is apparent, to which the term ebullition is usually applied.

Besides the views already taken of caloric, we have to notice other properties, which are embraced by the term *specific caloric*. When similar bodies, whose weights are equal, are exposed to the same degree of heat, they absorb precisely the same amount of caloric. But such is not the case, when substances differing in their nature, are exposed to the same elevated temperature. Balls of iron, copper, tin, &c., of equal weights, will indicate the same range of thermometric heat, after exposure in a fireplace for a given length of time; but experiments have shown, very conclusively, that the actual quantity of heat, in those balls, is not equal. Now, it is to these unequal quantities of heat, that the name, *specific caloric*, has been applied. It means the relative amount of heat, absorbed by different substances, by the time they evince a uniformity of temperature or sensible heat. The faculty, by reason of which bodies absorb these relative proportions, is called, the *capacity of bodies for caloric*; and analogy, as we shall presently show, confirms this statement.

The question has often been asked, what is specific caloric, and wherein does it consist? It is necessary, in order to solve this problem, to bear in mind, that whenever solid bodies, as metals, are heated very considerably, without being made fluid, two portions of caloric are received by them. One elevates the temperature and is cognisable by the thermometer, while the other simply enters into the solid, separates its particles by overcoming the force of cohesion, and thus augments its bulk. Philosophers have, therefore, very properly divided specific caloric into two distinct parts, each having a different function, viz. heating and dilatation. Hence, it really consists, first, of the *sensible* portion which is uniformly the same in different bodies, raised to the same temperature; and secondly, of the *latent* portion, which cannot be detected by the thermometer,



and is varied in quantity, by the difference of capacity, in each body, for caloric.

It is not easy to demonstrate the cause of this unequal capacity of bodies for caloric. That the particles of solids are detached from each other by the agency of heat, is certain from the most common mechanical operations, such as the expansion of tires for wheels; and aware, as we are, of the different densities of solids, we are at liberty to infer, that the repulsive power of heat is more energetic in some bodies, than in others, and of course, that the particles are more widely separated in one solid than in another. Now, as caloric is the grand repulsive power, it will be apparent to every reflecting mind, that it will accumulate more largely in bodies whose particles admit of the greatest degree of separation, than in those of the opposite character; and hence, we see a cause of diversity in the capacities of bodies for caloric, which is, perhaps, not inadequate.

Lavoisier performed some experiments with different kinds of wood, which seem to throw light on this question. He took several varieties of wood, and immersed them in water, in order to determine how much of that liquid each variety could absorb, the pieces being equal in bulk. The light and very porous woods, took in much more than those of a hard and compact texture, and a still greater quantity than the resinous woods, which, from their peculiar nature, exclude watery particles. In these cases, it is perfectly plain, that a difference of capacity for water was evinced, and that the diversity depended, chiefly, on the greater or less magnitude of the interspaces or pores, in the several species of wood. There is, evidently, a forcible analogy between the results of these experiments and the views already given, on the subject of capacity for caloric. In relation to the latter, however, there is superadded the powerful influence of affinity. Thus bodies which have a very strong affinity for caloric, will necessarily absorb it in large quantities. This position is strengthened, by the relation which many gases bear to charcoal. Numerous experiments have been performed, all proving that charcoal will take up, somehow or other, an immense bulk of gaseous matter. Thus, one volume of this solid will absorb ninety volumes of ammoniacal gas, eighty-five of muriatic acid gas, sixty-five of sulphurous acid gas, thirty-five of carbonic acid gas, &c., &c., &c.; and the gases thus absorbed, are retained so forcibly, that a heat of more than  $212^{\circ}$  is required to dislodge them. In these cases, however, it would seem, that the forces of the solids are not sufficiently capacious to contain such a bulk of gas, and that the latter must really be very much condensed. Here, too, it is plain that affinity performs an important share in effecting the results.



We have thus presented a view of the presumptive causes of the capacity for caloric; but how the *actual* quantity of caloric in bodies may be estimated, is a point not easily settled. We look in vain to the thermometer, or to any other instrument for measuring *sensible* heat, since the most accurate of these, has no cognisance of *latent* heat.

It has been seen, already, that when a body is heated, without changing from solid to fluid, one portion of the caloric that enters it, simply raises the temperature, while the other expands the particles. On cooling the same body, it loses, simultaneously, its sensible and latent heat; the reduction of temperature and diminution of bulk occur at the same moment. If we could seize upon both these portions of caloric, just as they are passing from the body, we might be able to form an estimate of its specific caloric.

Further, it has been apparent, that when bodies, after having been heated, are about to change from solid to fluid, or from fluid to gas, the heat which enters into them, does not, at the moment, raise their temperature, but simply serves the purpose of dilatation. Of all bodies that change their condition, none does so more readily than water. Ice melts, when its temperature is not above  $32^{\circ}$ ; and having ascertained the temperature of water used for this solution, we can determine the quantity of heat necessary to melt the ice. Every particle of caloric, thus communicated to it, goes, exclusively, to effect its change into the fluid state. Taking the quantity of ice, thus melted, we have a correct guide, by which to arrive at the quantum of caloric absorbed in the operation. It was on these principles, that Lavoisier invented his *calorimeter* or measurer of absolute heat, for an account of which, the reader is referred to that celebrated chemist's *Elementary Treatise*, a work which should be carefully perused by every student.

The capacity for caloric, varies in different conditions of the same body. Thus, ice, water, and steam, three states of the same substance, have not equal capacities for heat.

Two portions of the same body, in the same condition, have precisely the same capacity for caloric, as may be readily shown, by mixing a pound of water at  $50^{\circ}$ , with another pound at  $100^{\circ}$ . The temperature of the mixture, will be exactly  $75^{\circ}$ , or the mean heat of the two pounds.

It is often inquired, whether caloric is a material substance or a quality of matter; and there are facts which seem to favor both opinions. Hitherto, no bulk of caloric, however great, has been indicated by the most delicate balance, as, in the least degree, ponderable. While we behold its effects in expanding

iron, we do not witness any augmentation in weight, as an attendant of the expansion. It is alleged, however, notwithstanding this negative evidence, that more perfect balances may yet show that caloric is truly a ponderable substance.

That it is a mere quality, is thought to be highly probable, from analogy. Thus, it is affirmed, that a few grains of musk, while they impart their odor to a large volume of air, lose no appreciable weight. A rose, also, will diffuse its delightful fragrance through a room, and with the allowance of the evaporation of aqueous particles, is said to lose nothing of a ponderable character. What, then, peculiar to the musk and the rose, can thus affect our senses, if it be not some inherent quality of the material particles? And may not caloric, as a *property* of matter, produce all its effects, just as satisfactorily to the mind of the philosopher, as it could on the supposition of its material character?

A few observations on the sources of caloric, will be proper at this time. And as it was not the design of this work, to give a full history of caloric, the student will do well to consult the larger treatises, at his leisure; especially Thompson's book on *Heat and Electricity*.

Caloric is diffused throughout the wide domain of nature, and however different our sensations may be, in relation to various bodies, temperature may be said to belong to them all, relatively. Thus, in an apartment that is comfortably warm, the marble mantel will feel cold to the touch, while another body will indicate warmth. The mantel has less capacity for caloric, and is, at the same time, a better conductor, than many other bodies in the room. It takes the heat rapidly from my hand, and gives me the resulting sensation of cold.

*Friction* is one of the sources of heat. In this way, the savage tribes have, from time immemorial, been in the practice of lighting their fires. They rub two pieces of dry wood violently against each other, and sufficient heat is excited to cause inflammation. On the same principle, the working of ships' cables, has actually set the wood on fire, over which the cables moved, and various sorts of machinery have been acted on, in the same way.

If a small piece of phosphorus be rubbed on a board or slip of coarse paper, it is instantly inflamed. Mere friction will also cause some of the fulminating powders to explode. Exp.

If four or five grains of the salt called chlorate of potash, be mixed in a mortar, with two or three grains of Exp.

the flowers of sulphur, the friction of the pestle will cause the mixture to explode, light being at the same time evolved.

*Percussion* is another source of heat. The blacksmith, on this principle, kindles his fire. By quickly repeated strokes with his hammer, a piece of iron is made red hot, and thus affords the means of gaining his object. If a grain or two of

Exp. fulminating mercury be placed on an anvil, and struck smartly with a hammer, heat is evolved, and explosion ensues. The *percussion caps*, employed by sportsmen, are based on this very principle. A small bulk of detonating matter is placed in the brass or copper vessel, and the drawing of the trigger occasions the explosion.

*Compression.* The agency of this mode of exciting heat, is exemplified by the condensing syringe. At the hollow end of the piston, a portion of dry spunk is placed, and on forcing the piston quickly to the bottom of the barrel or cylinder, the air is so powerfully compressed, that sufficient heat is extricated, to fire the spunk. When glass syringes are employed, an evolution of light is apparent, at the instant of making the condensation. Some experiments of Thenard, make it probable, that the light and heat thus evolved, proceed, in part, at least, from the oleaginous matter which is attached to the piston, in order to make it work accurately. It is alleged, that when this oily substance is excluded, the spunk is not inflamed.

Heat is also elicited by the powerful compression of certain gases. This operation occasions condensation, and sensible heat seems to be a necessary result.

*Chemical action* is an abundant source of caloric,\* especially when bodies unite, which have a powerful affinity for each other. Thus, the substance called potassium, has a stronger affinity for

Exp. oxygen than any other body; and in its efforts to come in contact with this gaseous matter, it will burn on the surface of water.

If a few grains of chlorate of potash be mixed with an ounce or two of spirits of turpentine, previously warmed, and placed in a saucer, and a mixture of a half ounce of nitric acid and a drachm of oil of vitriol, also warmed, be suddenly thrown in, so great a degree of heat will be instantly excited, as to inflame the whole mass. As there is some risk in using the acids, it is advisable to put them in a bottle, fastened to one end of a stick, that is five or six feet in length. With

\* Heat is not always evolved by chemical action; on the contrary, intense cold is sometimes produced, as we shall have occasion to notice hereafter, when treating of salts that enter into the formation of freezing mixtures.

such a contrivance, they may be dashed upon the turpentine, without hazard.

The union of water and oil of vitriol, excites a heat, sufficient to inflame phosphorus, attached to the outer side of the vessel. Take an oil flask, and around its body, fasten some small pieces of phosphorus, by means of dry tow or cotton. Exp. Then pour in, successively, water and oil of vitriol, using four parts of the former to one of the latter. In a few moments, the phosphorus will be inflamed, and the tow or cotton kindled.

Pour cold nitric acid on some copper filings, and as chemical action goes on, a very considerable degree of heat will be evolved, sufficient, in some instances, to fracture the vessel. Exp.

*Combustion* is noticed as a distinct source of heat, and as it seems to be, in all cases, closely identified with chemical action, we shall not dwell on it here. In a separate chapter, an account is given of the various theories of combustion, embracing all that is important to the chemical student. The place allotted to this subject, seemed to be the only proper one; an earlier consideration would have anticipated, by necessity, several simple bodies, of which the novice could not be supposed to have a proper degree of knowledge.

It will be perceived, that the connection between caloric and the philosophy of the steam engine, is not brought into view in this chapter, and some may object to this apparent deficiency. But I hold, that the philosophy of steam engines has no more to do with chemistry, than the philosophy of watch making. Both appertain to the science of natural philosophy, in the strictest sense, and it will be found, that some of the best practical writers on chemistry are as silent on this subject as I have chosen to be.\*

*Electricity* and *galvanism* are powerful sources of heat, and shall be noticed, particularly, in this character, in the chapter appropriated to those subjects.

Of the *sun*, as the great source of caloric, a remark or two has been already made, and it is not needful to give the subject a special examination, in this place.

The terms *ignition* and *incandescence* are often met with in chemical books, and are to be regarded as synonymous. They are used to convey the idea of the accumulation of heat in bodies to such an extent, as to render them luminous.

*Cold*, which has been affirmed to be the mere abstraction of heat, is also a result of chemical action. Thus, where an im-

\* See Laugier, Reid, and others.

mense amount of caloric is absorbed by salts,\* in order to render them fluid, the whole of the caloric is said to be latent or insensible to the thermometer, and the falling of the mercury in this instrument, shows an evident reduction of temperature. Now, inasmuch as caloric is absorbed in great quantity, and as it passes into a state, not detectable by any measurer of heat, the conclusion is irresistible, that the caloric has entered into some sort of chemical union, which, although not easily explained, is not the less real. It would seem, at first view, that the more heat a body absorbs, the hotter it should be; but the reader will do well to recall to his mind the experiment, in which snow or ice was exposed to a fire, at such a distance, that the heat was  $172^{\circ}$ , and yet, when the solid was melted, the thermometer continued, for an instant, to stand at  $32^{\circ}$ . Here was an obvious absorption or taking in, by the snow, of  $140^{\circ}$  of heat, which, in consequence of a chemical union, was inappreciable by the thermometer. The same principle is applicable to all the varieties of *freezing* mixtures, and we shall seek in vain, for a better solution. Some further notice will be taken of this subject, when the different articles that are used for the formation of freezing mixtures, come under review.

The medical uses of caloric, and also of cold, or the abstraction of heat, claim a passing note; and the main reason for the remarks to be presented on this subject, is, to correct an error in regard to the *modus operandi* of *cold*. All agree that heat, whether internally or externally applied, in form of vapor, or water, or in any other way, stimulates the system and increases the action of the heart and arteries. On the simple principle of its repulsive power, its tendency to expand bodies subjected to its influence, one could hardly anticipate any other result from its application to the human body, than that which is constantly witnessed. These things being so, what must, on common sense principles, be the inevitable result of an abstraction of such an agent? Every one who is prepared to look at things just as they are, rather than to examine them through the spectacles of mysticism, will perceive, at once, that the abstraction of heat must give rise to effects of an opposite character to those which follow its direct application. It results, therefore, as a matter of necessity, that *cold* is a sedative; that is, an agent, (if that term be allowed to a mere negative,) which lessens the vital energies, and would ultimately destroy the principle of life, if not properly counteracted.

All admit that alcohol is a stimulus, because it is impossible to

\*As in freezing mixtures.



resist the daily evidence of our senses on this subject. An individual will keep his system up to the point of full toned excitement, by pursuing a regular course in the use of this poison. But suppose that some one who has power to restrain this lover of the unnatural stimulus, exerts his influence so as to stop, instantly, the wonted allowance, and to substitute cold water in its place. Here is a case of abstraction of a powerful stimulus, and look at the result. The man's whole frame indicates that a storm is gathering. His energies flag, and he is completely unnerved. Look at the drunkard who has delayed, for an hour or two, his morning dram, and see if his system cannot furnish evidence, that the want of the accustomed stimulus, operates as a sedative. Why, the unhappy creature has about him a temporary palsy; his hand shakes, his tongue falters, and he grasps the glass, as a drowning man would seize a straw. Look at him after he has swallowed the dose, and, in a few minutes, you discover, that the stimulus has had its accustomed effect; and the conviction is irresistible, that while alcohol is an unnatural stimulus to the drunkard, its abstraction is a tremendous sedative.\* It is precisely so with heat; its action is always to produce positive effects, while cold, operating in its negative character, always produces corresponding results. It is therefore a palpable contradiction, to say that *cold* is a *stimulus*.

---

### CHAPTER III.

#### Of Nomenclature.

LAVOISIER has observed, 'that every branch of physical science must consist of three things; the series of facts which are the objects of the science; the ideas which represented these facts; and the words, by which these ideas are expressed.' Hence, that acute philosopher labored to produce a perfect system of names; aware, not only, that the nomenclature of a science cannot be separated from the science itself, but that simple and intelligible names greatly facilitate study, and thus clear the way for advancement.

So long as barbarous and unmeaning terms were applied to the results of chemical action, the science, if it merited that

\* Let no one pervert these remarks into a defence of the use of alcohol. It is the mammoth curse of society, compared with which, famine or pestilence are not worth a name.

name, remained in the hands of, comparatively, few individuals. It had nothing about it to invite attention, but much to repel the beholder, and to awaken disgustful emotions in the mind of the cursory reader. The *colcothar*, *sal alembroth*, *pompholix*, and even the *Glauber* and *Epsom* salts, of olden time, are fair specimens of the enigmatical, crude, and undefined terms, with which the ancient books of chemistry abound. But the revolution effected by the Lavoisierian school, has excluded all this alchemistic jargon, and furnished a system of nomenclature, not only good in itself, but susceptible of easy application to all future discoveries. We now employ names that are designed, expressly, to convey correct notions of the composition of an article, if it be a compound; or to denote, with sufficient precision, some remarkable feature in the substance, if it be a simple. Thus, the study of chemistry has been greatly facilitated, and its importance has been estimated, in a corresponding ratio.

The names of elementary bodies are of little consequence, provided they serve to distinguish the one from the other. Of this nature, are *iodine*, *chlorine*, *bromine*, &c., in which are recognised, the *violet* color, the *yellowish green*, and the *fatid* smell, peculiar to these substances.

With regard to compounds, the case is otherwise. It is necessary to employ names indicative of the actual composition of these bodies, however complex they may be. Thus, the term *oxyde*, refers to the union of oxygen with some base, as iron; and hence the oxyde of iron. Sometimes the compound resulting from the union of oxygen and a base, is called an *acid*; thus, oxygen and sulphur form the sulphuric acid. So also iodine, chlorine, and bromine, united to bases, form *iodides*, *chlorides*, and *bromides*, which are analogous to the oxydes, in many respects. Thus we have *iodide of mercury*, *chloride of potassium*, &c. But when these simple bodies are combined with oxygen, or hydrogen, they give rise to acids, and hence the *iodic*, *chloric*, and *bromic*, as well as the *hydriodic*, *hydrochloric*, and *hydrobromic acids*.

The terms oxyde, iodide, chloride, and bromide, denote compounds of the several elementary substances, oxygen, iodine, &c., with certain bases; but as the relative proportions of these elements is not uniformly the same in all their compounds, it is necessary that terms be added which may denote with sufficient precision, the actual proportions. Hence, the introduction, by Dr Thomson, of the prefixes, *proto*, *deuto*, *trito*, and *per*. Thus, the *protoxyde* of iron, the *proto-chloride of mercury*, and the *proto-iodide* of lead, announce compounds in which oxygen, chlorine, and iodine are found in the first or lowest proportion. If these

simple substances have but two proportions of combination, the first is called a *protoxyde* or proto-chloride and the other, a *peroxyde* or a per-chloride. If four varieties of combination obtain, then we have the protoxyde, deutoxyde, tritoxide, and peroxyde, as also, the proto-chloride, deuto-chloride, trito-chloride, and per-chloride.

Now it may be affirmed of these substances, as a general rule, that they have no action on vegetable blue colors, whereas acids generally change those colors to red. The acids are usually distinguished from each other, by prefixing the radical of each. Thus, the acid of sulphur and the acid of phosphorus, are called *sulphuric acid* and *phosphoric acid*. The termination *ic* is applied to an acid that contains oxygen, if it be present at all, in the largest proportion. When the name of the acid ends with *ous*, we know that the radical has combined with a less quantity of oxygen, and hence we have the *sulphurous* and *phosphorous* acids. These varieties were, for a long while, the only ones in use, but as discovery progressed, there appeared to be a propriety in multiplying terms to meet the exigencies of the case. Thus, an acid was discovered by Guy Lussac, which was not, strictly, either sulphuric or sulphurous, but something between the two; and another was investigated which appeared to be a grade below the sulphurous acid. To meet these cases, the word *hypo*, from the Greek, meaning *under*, was introduced, and is now in general use. So we have sulphuric acid, which is sulphur in the highest state of oxygenation; *hypo*-sulphuric acid, a compound of sulphur with a less proportion of oxygen; sulphurous acid, in which the quantity of oxygen is still less; and *hypo*-sulphurous acid, in which there is the smallest proportion of oxygen, with which sulphur can unite, to form an acid compound.

Formerly it was supposed, that all acids necessarily contained oxygen, as an *acidifying principle*; but as the union of chlorine, iodine, &c., with hydrogen, gives rise to acids, we now distinguish between *oxacids*, and *hydracids*. Thus, sulphur, chlorine, iodine, bromine, &c., in union with oxygen, form sulphuric, chloric, iodic, and bromic acids; while the same articles joined to hydrogen, and entirely apart from the influence of oxygen, form hydro-sulphuric, hydrochloric, hydriodic and hydrobromic acids. Whenever the term *hydro* is prefixed to an acid, we are to regard it as an abbreviation of hydrogen, which substance is a constituent part of the acid.\*

On the subject of an acidifying principle, I have given my

\* The term *hydro* has been, sometimes, used as a prefix, to signify water.

views at length, some years ago.\* It may not be amiss, however, to state, in this place, that the advances which chemical science is constantly making, have confirmed my early opinions on this point. I repeat, that the term acidifying principle is utterly unphilosophical, not only as applied to oxygen, but to hydrogen, and to every agent which may be supposed to exert an influence in developing acid properties. Every result in nature or produced by art, is a relative effect, and every item concerned, remotely or directly, in the accomplishment of the end, is essential to that end. Hence, I insist, that if an acid be discovered, which shall contain fifty component parts, all of which are requisite in the formation of the compound, the only characteristic of which is acidity, I may affirm with equal propriety of any one, as of any other, of its constituents, that this or that is the acidifying principle. Abstract from the compound either of its parts, and you destroy the peculiar, distinctive character of the acid.

The *green copperas* of the shops is generally known, as a *sulphate of iron*; that is, a combination of sulphuric acid and iron. But in view of the various degrees of oxydation, as expressed by the prefixes *proto*, *deuto*, *trito*, and *per*, chemical accuracy demands a more precise nomenclature, than that which common usage allows. To say that copperas is composed of sulphuric acid and *oxyde* of iron, is not sufficient, because, as we have already seen, there are varieties in the oxyde of iron. It is therefore necessary to adopt a name which will announce, precisely, the degree of oxydation in which the metal is found united to the acid. This distinction is not merely a matter of chemical accuracy, but one of practical importance, as will appear in the sequel. Now the iron in green copperas, is in the state of a protoxyde, or in the first degree of oxydation, and, therefore, instead of using the term sulphate of iron, we call it proto-sulphate of iron, or sulphate of the protoxyde of iron. Green copperas is a term as vague and indefinite as man can devise, but either of the two names last mentioned, is so significant, that nothing more can be desired, to remove obscurity.

If sulphurous acid combine with a base, as iron, the compound is not called a *sulphate* but a *sulphite*. The compounds of hypo-sulphuric and hypo-sulphurous acids with bases, are denominated hypo-sulphates and hypo-sulphites, and the same rule applies to all the varieties of phosphoric, nitric, and other acids.

\* See Memoirs of Columbian Chemical Society, and New York Medical Repository, published in 1812 and 13.

*Binary* salts or compounds are formed of two ingredients; thus hydrogen and oxygen constitute water. In *ternary* compounds, three articles are present; thus alum is composed, ordinarily, of sulphuric acid, alumine, and potash. *Quaternary* compounds, (they are very rare,) consist of four ingredients.

The term *neutral* is applied to a compound, in which the marked characters of each ingredient are lost. If an acid and a base, as for instance, nitric acid and potash, be united in due proportions, the properties of both will disappear, and the salt thus formed, is said to be *neutral*. Thus, sulphate of soda, (Glauber's salts,) may be formed by uniting sulphuric acid and soda, in certain quantities, and the product is a neutral salt, possessing none of the characteristic features of either of its constituents. In some compounds, the acid is said to be in excess; that is, more acid is present than serves simply to form a neutral salt. This excess of acid is detected, by its changing litmus paper from blue to red. The salts thus constituted, have been called *super* salts, to distinguish them from *sub* salts in which the acid being deficient, the base is in excess. Thus, the books have the *super*-carbonate of soda, and *sub*-carbonate of soda; the former containing an excess of acid, the latter an excess of base, or alkali. These terms are now very much laid aside, because the doctrine of definite proportions has given us substitutes more exact, and less calculated to mislead. The borax of the shops was formerly called a *sub*-borate of soda, because it exhibits alkaline properties, as is apparent from its action in changing turmeric paper, from yellow to brown. But it is now certain, that this salt contains two equivalents or proportions of boracic acid, to one of soda, and it is called a *bi*-borate of soda. This prefix *bi*, is an abbreviation of the latin *bis*, signifying twice or double. The common *cream of tartar* and *soluble tartar* of the shops, have been called *super*-tartrate, and tartrate of potash. The latter is a *neutral* salt, having neither acid nor base in excess; but the former contains two proportions of tartaric acid to one of potash, and is, therefore, termed a *bi*-tartrate.

The term *sesqui*, has found its way into our systems of chemistry, and was introduced by Dr Thomson. It seems to have pretty much the same relation to salts, that *hypo* bears to acids. Thus, where there are two acknowledged compound salts, in which the acid is of the same kind, but differing in proportion, as one to two, if a *sesqui* salt be discovered, containing the same constituents, the acid will be in the proportion of one and a half. I do not think, however, that this term is likely to be at all beneficial to the interests of chemical science.

The terms *anhydrous* and *hydrates* frequently occur in modern



treatises, and they are not without use. The former denotes a substance, either simple or compound, which does not contain water; the latter always denotes the presence of that liquid, and of course, they may be regarded as opposite terms.

---

## CHAPTER IV.

### Of Specific Gravity.

As we shall have frequent occasion to speak of the specific gravity of various bodies, it is proper to notice the subject in this part of the work. By specific gravity, we mean the relative weight of bodies. Archimedes was the first to teach, 'that a body immersed in a fluid, loses as much in weight, as the weight of an equal volume of the fluid;' and he determined, by means of this principle, how much alloy an artist had fraudulently added to a crown, which king Hiero had ordered to be made of pure gold. He discovered the solution of this important problem, while in the act of bathing; and it is said to have elated him so much, that he ran from the bath, undressed, to his home, crying along the streets, 'I have found it, I have found it.'

The standard of comparison, in calculating the specific gravity of most bodies, is water, and it is reckoned at one or one thousand.

A hydrostatic balance or good pair of scales, (as represented in Fig. 2d, Plate I,) and a glass tumbler, are all the instruments necessary, in order to determine the specific gravity of a solid body, that is heavier than water. It is necessary to have one of the scales fastened within a short distance of the beam, and to fix a hair or fine thread to the bottom of it, for the purpose of suspending the body to be weighed. This being done, first weigh the substance in the air and note down its weight accurately; next it is to be weighed in water, by simply letting it pass into the tumbler nearly filled with that fluid, and the exact weight is to be taken. The difference in the result will express the weight of a quantity of water equal in bulk to the solid, whose specific gravity is to be determined, and the following proportion will give us its specific gravity in relation to the water; as the weight of the water, equal in bulk to that of the solid, is to the weight of the solid itself, so is the specific gravity of water to the specific gravity of the solid. Thus, suppose the solid to weigh one hundred grains in air, and sixty grains in water, the difference between these will be forty. If we divide the weight in air, viz. 100. by the difference between the weight

in air and in water, viz. 40, we shall have 2.5 as the specific gravity of the solid, compared with that of water.

If we want the specific gravity of a solid, lighter than water, a more complicated process will be necessary. We must attach to the light body, by means of a fine thread, another body, sufficiently heavy to sink both in water; having previously taken the weight of the heavier solid in water, and of each in air, we are next to weigh them jointly in water, and from the difference between their weight in water and in air, we must subtract the difference between the weight of the heavy solid in air and its weight in water. The remainder will indicate the weight of a quantity of water equal in bulk to the light body, and we can thus find its specific gravity in the way directed above. Thus, let the weight, in air, of the lighter solid be 10, and of the heavier solid 20; the weight of the latter in water, 18, and of both together, in the same fluid, 7. Then from the joint weight of each solid in air, viz. 30, deduct the weight of both in water, viz. 7, and we have 23; from this deduct the difference between the weight of the heavier solid in air and in water, viz. 2, and we have 21, which expresses the weight of a quantity of water equal in bulk to the light solid, and the following proportion will give us its specific gravity.

As 21 is to 10, so is 1 to 0.47619, or the specific gravity of the lighter solid.

A hydrostatic balance is not always at hand, and the want of it may be supplied by the following expedient. Let the solid body be carefully weighed and then placed in a vessel full of water, ascertaining at the same time, the weight of the whole. The solid will be found to displace a quantity of water equal in bulk to its own; let the vessel be again weighed, having first taken out the solid, or put an equal weight into the opposite scale. The difference between the present weight of the vessel and its former weight, will express the weight of a quantity of water equal in bulk to the solid body, from which, by the same proportion, as in the former instances, we can estimate the specific gravity of the solid body. Thus, if the vessel, when full of water, weighed 1000, and after the displacement of a portion by means of the solid body, it weighed 900 grains, it is obvious that 100 grains of water have been displaced by the introduction of the solid body. If this latter weighed 300 grains in air, then the following proportions will show its specific gravity;

As 100 is to 300, so is 1 to 3.

If the solid body be soluble in water, we must use another fluid, as wine, or alcohol, which do not act on it. Then, supposing the specific gravity of the body, thus ascertained, to be

1.8, and the specific gravity of the fluid used in the process to be 1.2, we say,

As 1.2 is to 1.8, so is 1 to 1.5, or the real specific gravity of the solid.

For other modes of finding the specific gravity of solids, we refer the reader to the books of natural philosophy.

To determine the specific gravity of liquids, we take a bottle of a given weight, fill it with distilled water and weigh it carefully; then pour out the water, and after the bottle is quite dry, fill it with the liquid to be examined. The following proportion will present us with the specific gravity. As the weight of the distilled water is to the weight of the other liquid, so is 1 to the specific gravity sought. Thus, let the distilled water weigh 150, and the liquid under examination 200; then,

As 150 is to 200, so is 1 to 1.333.

An instrument, called a *thousand grain bottle* has been used advantageously for the same end. It is accompanied by a weight, which is an exact counterpoise for it, when filled with distilled water at 60°, Fahr. This instrument does not require the aid of computation, but is simply filled with the fluid to be examined, and placed in a scale of the balance, while its counterpoise is placed in the other. If the contained fluid be lighter than water, it will appear deficient in weight, and as many grains must be added to the scale that contains it, as may be sufficient to restore the balance. But should the fluid be heavier than water, the bottle will preponderate and weights must be put into the opposite scale, the amount of which must be added to the weight of the standard, viz. the water. Thus, suppose the bottle to be filled with sulphuric ether; at least 739 grains must be placed in the same scale to restore the balance, and consequently, the specific gravity of the ether must be 0.739. If the bottle contain sulphuric acid, 875 grains must be added to the opposite scale in order to balance the acid, so that 875 must be added to the weight of the standard, viz. 1000; thus making the specific gravity of sulphuric acid to be 1.875.

Areometers or hydrometers are convenient instruments for ascertaining the specific gravities of liquids. They are of various form and size, differing in these respects, according to the peculiar notions of their inventors. They consist, generally, of a long, straight, graduated stem, on which numbers are marked at the points to which the instrument sinks in liquids of the specific gravities denoted by these points. Thus, in distilled water it will sink to 1, and in nitric acid to 1.5.

In taking the specific gravity of gases, atmospheric air is employed as a standard, and represented as 1. We compare the

weight of equal bulks of them, and of the standard of comparison, and so arrive at the desired conclusion.

For this purpose, a glass flask or globe, (Fig. 11th, Plate 2d,) provided with a stop-cock, is accurately weighed, and attached to an air pump or exhausting syringe, which is worked in the usual manner; and when the gas to be examined, has no action on atmospheric air, it is not needful to exhaust it to a very great degree. The stop-cock fixed to the flask is then turned, and again weighed to ascertain the quantity of air extracted. We next screw it to a jar containing the gas to be tried, placed over a pneumatic tub, and on turning the stop-cock, the flask or globe is soon filled, if the jar be depressed in the water. The stop-cock is then to be closed and the vessel again weighed. Suppose the flask to weigh 570 grains when full of common air, and 560 when exhausted, then it is plain that 10 grains have been withdrawn. If it weigh 580 grains when filled with the gas under examination, it is equally certain that 20 grains of gas have been introduced. The flask contained 10 grains of common air in the first instance, and now holds 20 grains of a particular gas, and in both instances was filled. It follows, conclusively, that the specific gravity of the gas under trial is exactly double that of common air.

When the gas, whose specific gravity we desire to know, acts chemically on common air, the latter must be extracted as perfectly as possible by repeated exhaustions, filling it after each with some gas which is not affected by the other, and then proceeding in the usual manner.

It is to be understood, that whenever the specific gravity of a solid, liquid, or gas is mentioned, the amount specified is to be viewed relatively to the specific gravity of pure water or atmospheric air.

To show that atmospheric pressure has an influence on the relative weight of bodies, the experiment with the air-pump and guinea and feather apparatus, may be advantageously exhibited. In lieu of the guinea and feather, we commonly use a piece of lead and another of cork, of equal size. If these be dropped from a height, in open air, the lead falls to the ground before the cork. But in the experiment alluded Exp, to, the cork arrives at the ground, quite as soon as the lead.

## PART II.

OF SIMPLE OR UNDECOMPOSED SUBSTANCES.

### CLASS FIRST.

OF SIMPLE SUBSTANCES NOT METALIC, AND THEIR COMBINATIONS  
WITH EACH OTHER.

---

#### CHAPTER I.

##### Oxygen.

THE first undecomposed or simple substance to be noticed is oxygen. In combination with caloric, it is denominated oxygen gas. It was discovered by Dr. Priestley, in 1774, and by Scheele in a year or two after, without any previous knowledge of what Priestley had done. The appellations given to it have been various. Thus, it was called *dephlogisticated* air by Priestley; by which term he meant, air freed from phlogiston. Scheele called it *empyreal* air, or the air of fire; and Condorcet named it *vital air*, or the pabulum of life. The term by which it is now known was given by Lavoisier, on the supposition that it was the sole cause of acidity; being derived from two Greek words which signify, to *generate acid*.

This simple substance is more widely diffused throughout nature than any other of that class of articles. It enters the composition of vegetable substances, animal solids, and a multitude of mineral productions. It forms a part of most liquid substances, as water and the acids, and enters into the formation of many gases. It constitutes a fifth part of the atmospheric air, and eight-ninths of the waters of the globe. In its gaseous form it is transparent and colorless, and pre-eminently distin-



guished by its power of maintaining combustion and respiration. It can unite with all the elementary substances, and with many of them in several different proportions. When it combines with the non-metallic bodies, the resulting compounds have acid properties in most cases. With the different classes of metals, it forms alkalies, earths, and the common metallic oxydes.

The equivalent of oxygen, by weight, is 8, compared with hydrogen as 1. By volume, its equivalent is a *half measure*,  $\equiv$ , compared with hydrogen as a *whole*,  $\square$ . Thus, 8 of oxygen combine with one of hydrogen, by weight, to form water; or a whole measure of hydrogen unites with a half measure of oxygen to form the same fluid. The specific gravity of oxygen is 1.111, compared with common air as 1, at the temperature of 60° Fahrenheit, and 30 of the barometer. One hundred cubic inches of this gas weigh 33.888 grains, while the same quantity of common air weighs 30.5 grains. This relative weight of 100 cubic inches of oxygen or other gases with the same quantity of common air, affords a satisfactory mode of determining their specific gravity. Thus,

As 30.5 is to 33.888, so is 1 to 1.111, or the specific gravity of oxygen gas.

This gas is very sparingly absorbed by water, and evolves both light and heat when suddenly compressed. It has neither taste nor smell, and is a non-conductor of electricity. It is the most perfect negative electric that we possess, always appearing at the positive pole of the galvanic battery.

The substances from which oxygen gas is prepared, are, the common saltpetre of the shops, the black oxyde of manganese, red lead, red oxyde of mercury, and the chlorate of potash. The last article yields by far the purest gas, but it is too expensive for laboratory use, when a large quantity is wanted. The black oxyde of manganese is the cheapest substance for making oxygen, but as it is not often obtained of good quality, it frequently disappoints the operator. I have never failed with the saltpetre, (or nitrate of potash,) and prefer it decidedly for all experimental uses.

We may obtain oxygen from saltpetre by placing it in an iron bottle, and exposing it to a red heat; and in the same way it is often procured from the black oxyde of manganese. The iron bottles can be had of the apothecary, who receives his quicksilver in them, and who is always willing to dispose of them at a moderate price, as they cease to be of much value to him after they have been emptied of their contents. The neck of such a bottle has a large screw adapted to it, so as to prevent effectually the escape of the mercury when the bottle is accidentally in-

verted. This screw being removed, a gun-barrel must be fitted to the neck, by a smith, so as to take the place of the screw, and then, being curved so as to form nearly a right angle with the bottle, it is fitted to convey gases, either directly into gas holders, or into other tubes which communicate with suitable receivers. The saltpetre is to be coarsely powdered, and introduced into the iron bottle, and then the gun-barrel is to be screwed in. About a pound or a pound and a half of the salt will be as much as ought to be put into the bottle at once, and it will yield a considerable quantity of gas. The bottle may then be placed in an open coal fire, or in a small furnace, and heated to redness. If it is desirable to convey the gas to a considerable distance, a small leaden pipe may be attached to the gun-barrel and secured by a common lute of flour and water with some white of eggs and a piece of linen or cotton cloth. Thus, the gas may be conveyed into glass receivers, inverted over water on the pneumatic tub, (as in Fig. 1st, Plate 2d.) or, on the same principle, into wooden or tin receivers placed within the tub.

The same process, precisely, will answer for making this gas from manganese. In both cases, however, the gas must be tried before we collect it in the receivers, in order to be sure that pure gas is passing over. The nitre or saltpetre consists of an acid called *nitric*, and a base, viz. *potash*; and the first effect of the heat is to melt the compound. As it contains a good deal of what is termed water of crystallisation, it boils up considerably, and when a very small bottle is used, there is danger of explosion from the choking up of the neck. This risk is obviated by using the large iron bottle already noticed. Soon after the salt is melted, the work of decomposition begins. The nitric acid is resolved into oxygen and nitrogen, which are its elementary principles, or probably, as some suppose, it is decomposed into oxygen and nitrous acid. The oxygen is evolved, and the nitrous acid combines with the potash to form the nitrite of potash. But if the heat be continued after the oxygen has escaped, then the nitrous acid, in combination with the potash, will be entirely decomposed, and the gaseous elements, passing over, leave behind nothing but peroxyde of potassium, or the metallic base of the potash, combined with the largest amount of oxygen with which it can unite. This peroxyde of potassium is found to yield oxygen, instantly, when water is thrown upon it, and has been proposed as a good substance for making oxygen gas, when small portions are wanted.

It often happens that when decomposition has commenced, either nitrogen or nitrous gas is evolved, and on applying a lighted taper to the mouth of the gun-barrel, it is instantly extin-

guished. But so soon as pure oxygen comes over, the taper, instead of being extinguished, will burn with increased brilliancy, and this is the time for collecting the gas in the proper receivers.

If we use the black oxyde of manganese, the same precaution is requisite to determine when pure gas is coming over. The first portions are common air and carbonic acid, the latter of which arises from carbonate of lime, accidentally present, and which renders the product unfit for supporting combustion. In a few minutes, however, we shall find that pure oxygen is given out, if the manganese be of good quality. Red lead, (red oxyde of lead,) and the red oxyde of mercury, as well as the chlorate of potash, can be made to yield oxygen gas, by a contrivance similar to that in Plate 2d, Fig. 1st. The articles are to be introduced into the retort, *B*, and gradually heated by the lamp below, until decomposition is effected. In the case of the two oxydes, the heat simply drives off the oxygen gas; from the oxyde of mercury all the oxygen is expelled, but only a part is driven off from the oxyde of lead. In all these cases, the retorts should be of light green bottle-glass, because they are not so easily softened as those made of flint-glass; the temperature required to expel the gas, almost melts the white retorts, owing to the large quantity of oxyde of lead which they contain. In addition to this dry way of obtaining oxygen gas, it may be procured from the black oxyde of manganese, by adding to it rather more than its own weight of sulphuric acid. Every 44 parts of the black oxyde require 49 of acid, for this purpose; and a much less degree of heat is requisite than when the dry method is pursued. In this case, a tubulated retort, (Fig. 2d, Plate 2d,) will answer better than the plain retort represented in Fig. 1st. The acid should be poured through a funnel passed into the upper aperture of the retort, usually called the tubulure. Then, through a funnel perfectly dry, the powdered manganese should be introduced, and the whole should be well shaken before heat is applied.

It is proper to remark, that a chafing-dish, or chauffer, as some writers call it, will be found preferable to an Argand's lamp for heating retorts. When the coals are fully ignited, they give a more uniform heat and will be found to answer a very good purpose.

In preparing oxygen and other gases in retorts of glass, great care must be taken to maintain a constant and uniform heat, and to remove the retort when the gas ceases to come over. By neglecting this precaution, a vacuum will be formed in the body of the retort, and the water from the cistern will rush in

and fracture the vessel, if no other accident occur. The water in these cases, is said by chemists to be *absorbed*; and as the cause of the accident is obvious, a little care will readily prevent it. If the retort has been exposed to a high degree of heat, and the mass within is quite hard and solid, the heat must not be withdrawn at once, but gradually; otherwise the cold air playing on the external surface, while the heated matter within keeps the internal surface hot, will be very apt to occasion a fracture. It is on this principle, precisely, that a tumbler is broken in cold weather, by suddenly pouring hot water into it.

We shall now offer a few remarks, explanatory of the formation of this gas from the various substances already spoken of.

In forming oxygen from saltpetre, we have to bear in mind, as already hinted, that it is a compound of nitric acid and potash, and that this combination is broken up, by the agency of heat. The following diagram will show the composition of the saltpetre, before it is exposed to heat, and the products arising from decomposition. The equivalent number of nitric acid is 54, and that of potash 48, because these substances always combine in those proportions.

54 nitric acid,	{ oxygen,	8—8 oxygen.
48 potash,	{ nitrous acid, 46}	94 nitrite of potash.
	48}	
<hr/> 102 nitrate of potash.		<hr/> 102

Thus, we see that the nitric acid, being decomposed into oxygen and nitrous acid, the former escapes in form of gas, to be collected in proper receivers, while the latter joins the potash and forms nitrite of potash. A pound of saltpetre will furnish about 1200 cubic inches of gas.

To understand the theory of the formation of this gas from the black oxyde of manganese, we must recollect that there are three oxydes of this metal, of which the following is probably the correct atomic view.

	Manganese. Oxygen.			
Protoxyde,	28	+	8	= 36
Peroxyde,	28	+	16	= 44
	Protoxyde. Peroxyde.			
Deutoxyde,	36	+	44	= 80

Two equivalents of the peroxyde will, of course, be 88; which, on the application of heat, will give out 8 parts of oxygen, leaving 80 of the deutoxyde of manganese in the bottle or retort;



but when we use sulphuric acid, and apply a more gentle heat, every 44 parts, or one equivalent of the peroxyde, (black,) will give 8 parts of oxygen, and 36 parts, (or one equivalent,) of protoxyde will remain, combined with 40, (one equivalent,) of dry sulphuric acid. Every equivalent of the black oxyde, equal to 44 grains, loses, if pure, 4 grains of oxygen, or nearly 12 cubic inches, by the dry mode; and in this way, one ounce will yield about 128 cubic inches of gas. This is the most convenient method with the manganese.

In forming oxygen from red lead, we take what is called the deutoxyde, which is a compound of the protoxyde and peroxyde of lead. Thus, the

	Lead.	Oxygen.	
Protoxyde is composed of	104	+	8 = 112
Peroxyde,	104	+	16 = 120
	Protoxyde.		Peroxyde.
Deutoxyde,	112	+	120 = 232

Now the effect of heat is to expel from this deutoxyde, = 232, precisely 8 parts of oxygen, leaving 224, or two equivalents of protoxyde in the retort.

If we make oxygen from the red oxyde of mercury, (the peroxyde,) we must recollect that it is a compound, of mercury 200, and oxygen 16, = 216; or one equivalent of metal and two of oxygen. The heat expels all the oxygen, and at the same time volatilises the mercury, which is again made fluid by condensation in some part of the apparatus. Owing to the volatility of the mercury, we must be careful that none of the fixtures with which the vaporised metal comes in contact, are made of any metallic substance.

We notice last, the phenomena attending the forming of oxygen from the chlorate of potash, a salt formerly known by the name of oxymuriate of potash. For making the gas from this substance, a green glass retort should be selected, if it can be had. I have, however, used the white glass, but have always witnessed a partial fusion, and the round bottom of the retort is often drawn down almost to a point. I prefer a tubulated retort in this process, because the salt can be all deposited directly in the bottom of the vessel, so that the heat will act on the whole at once. In a few minutes, the salt is made quite fluid, and its bulk is enlarged. We should be careful, in order to avoid danger, not to introduce more than one ounce of the salt into a half pint retort, at one operation.

The chlorate of potash contains more than a third of its weight of oxygen, and all this is expelled by heat. Every 124 grains yield 112 cubic inches of pure oxygen gas.



To understand the process, it is proper to say, that 76 is the combining or equivalent number of chloric acid, which united to 48, the equivalent of potash, make 124, or one equivalent of chlorate of potash. The chloric acid is composed of oxygen and another simple body called chlorine; while potash is a compound of oxygen and a metal, named potassium. Thus,

$$\begin{array}{rcl}
 76 \text{ chloric acid} & = & 36 \text{ chlorine} \quad + \quad 40 \text{ oxygen.} \\
 48 \text{ potash} & = & 40 \text{ potassium} \quad + \quad 8 \text{ oxygen.} \\
 \hline
 124 \text{ chlorate of } & \} & 76 \text{ chloride of } \} \quad 48 \text{ oxygen.} \\
 \text{— potash.} & \} & \text{— potassium.} \} \quad \text{—}
 \end{array}$$

Of course, all the oxygen contained in the chloric acid as well as in the potash, is driven off, amounting to 48 parts; while the chlorine of the acid and the potassium of the potash, unite to form 76 parts of chloride of potassium. The latter remains in the retort, and being soluble in water, is easily removed.

In addition to the above modes for preparing oxygen gas, we may obtain it, by means of the green leaves of vegetables, immersed in an inverted jar full of water, and placed in the sunshine. The appearance of oxygen gas, in this process, is variously explained. Some say, that the water is actually decomposed, its hydrogen taken in by the plant, while the oxygen is liberated. Others affirm, that the oxygen is obtained by the decomposition of carbonic acid, which is present, the carbon being taken up by the plant, and the oxygen evolved. Dr. Manners, of New Jersey, performed some experiments in 1812, to prove this theory.\* When lime water was employed, instead of pure water, no oxygen was obtained; the same thing occurred, if the water was boiled before the leaves were placed in it. Now, in the first case, all the carbonic acid was seized by the lime; and in the latter, it was expelled from the water by heat.

In addition to the properties of oxygen already named, we may affirm, that it has neither acid nor alkaline qualities, for it does not redden, nor turn green the vegetable blue colors, nor does it evince any tendency to unite with acids or alkalies. For most of the simple bodies it has a strong attraction, and is susceptible of union with all of them. The act of combining with oxygen, is sometimes called *oxydation* and also *oxygenation*, and the bodies resulting from this union, are said to be oxydised and oxygenated; they are also called *oxydes*, and some of them *acids*.

Many bodies acquire oxygen from the air, if moisture be present. Thus, iron soon passes into *rust*, and green copperas is

\* See Memoirs of the Columbian Chemical Society.

formed by exposing the native sulphuret of iron, in the same manner. The sulphur acquires enough oxygen to form sulphuric acid, while the iron is oxydated, and the acid and oxyde join, to form the copperas.

When acids containing oxygen are poured, in a concentrated form, on substances that have a great affinity for this element, as metals and inflammable bodies generally, they rapidly give out their oxygen, especially if heat be employed. Thus, mercury is speedily oxydated by nitric acid, without much heat; and if this metal be boiled in sulphuric acid, the same thing takes place.

The phenomena of oxydation are very variable. Thus, in the rusting of iron, no light is evolved, nor do we perceive any escape of heat; and the slower the process, the less perceptible will these results be. In active, quick combustion, as in the burning of what are called *combustible* bodies, both light and heat are obvious products, and yet the ordinary instances of combustion are really nothing more than cases of rapid oxydation.

The burning of bodies in common atmospheric air, is dependent on the oxygen which enters its composition, and just in proportion to the quantity of oxygen present, will be the activity of the process. Bodies that burn well in the open air, burn still better in oxygen gas. Hence a piece of wood on which there is a single spark, bursts into flame, when immersed in this gas. Lighted charcoal, also, throws out vivid scintillations when transferred from a jar of common air to another of oxygen gas. Phosphorus burns smartly in the air of the atmosphere, but it emits a light when burning in oxygen, which the eye cannot tolerate. So, also, an iron wire of the finest kind, and heated to redness, burns rapidly in this gas, although this metal is not, usually, ranked among inflammables. If we tie some tow round the end of a piece of fine iron wire, then dip it in melted sulphur, and immediately after kindling, immerse it in a jar of oxygen gas, the wire will be burnt, throwing out brilliant sparks on all sides. The glass vessel in which this experiment is conducted, is often fractured, by the red hot particles of iron;\* but this accident may be frequently obviated, by previously pouring water into the vessel to the depth of an inch. The mode of thus burning iron wire, may be seen in Fig. 10th, Plate 2d.

\* Such is the velocity with which the particles of iron are driven against the glass, that I have several times seen them actually indented in the substance of the bottle.

We may readily show the superiority of oxygen over common air, as a supporter of combustion, by having two jars, containing these gaseous matters, separately. Into the latter,

Exp. lower a lighted candle, and it will burn as in a common room; then transfer it to the other jar, and the combustion will be vastly more brilliant. This property of oxygen is further evinced by the following experiment; having a jar of oxygen gas at hand and a lighted candle, extinguish the latter, and instantly dip it into the jar. You see it immediately rekindled, with a slight explosion. Extinguish it again, and again immerse it in the gas, and it is speedily lighted. This experiment may be repeated several times with the same jar of gas.

If we take a little metallic arsenic, moisten it with spirits of turpentine, then fire it, and quickly immerse it in a jar of oxygen gas, the arsenic will burn in a very beautiful manner.

Exp. So also if we enclose a piece of phosphorus in some turnings of zinc, so as to make a ball, and having set the phosphorus on fire, quickly dip the ball into a vessel of oxygen gas, it will burn with a beautiful white light.

In all cases of combustion, where oxygen is the supporter, that process will continue no longer than free oxygen is present, because the very existence of combustion depends on the combination of this gas with the inflammable body, and hence arose the term, *supporter of combustion*. That the gas actually disappears, is strikingly obvious in the burning of phosphorus over water, by the ascent of the liquid, which is forced up into the vessel, by atmospheric pressure.

Not only does the oxygen disappear, and wholly lose its properties, but the combustible body is also changed. This will no longer burn, even in the purest oxygen; and it is called an oxydised body, partly to indicate that it has already been subject to the influence of oxygen. Thus, in the burning of our candles, the tallow and wick are combustible matters, and the oxygen of the air, maintains the combustion. Now, although all these are changed and seem to be destroyed, they have only entered into new states. The oxygen has combined with the carbonaceous matter of these combustibles, and carbonic acid has been produced. And, however opposed to vulgar notions, it is nevertheless true, and susceptible of demonstration, that the body burnt always acquires an increase of weight, which is proportioned to the quantity of oxygen consumed in the process of combustion.

But while oxygen is essential to the ordinary cases of combustion, it is absolutely indispensable in the process of respira-

tion. It is not possible for an animal to live, for even a short period, in an atmosphere entirely deprived of oxygen. An air that will instantly extinguish a lighted candle, will very speedily put out the lamp of animal life.

The salutary effect of oxygen on the blood, is well known to physiologists. Its fresh supply in the lungs, in every act of respiration, is essential to the healthful actions of life. In those cases of malformation, in which the blood passes, for the most part, from the right to the left side of the heart, without experiencing this beneficial change, the unhappy sufferers linger out a miserable existence, and end their days before attaining to the age of maturity. Their countenances are livid, and hence, they are sometimes called *blue children*.\*

The next notice to be taken of oxygen gas, will embrace its medical uses. The time has been, when oxygen had considerable reputation as a remedial agent. It is to be confessed, however, that its employment was for the most part of an empirical character, even in the hands of regular practitioners. I doubt not that there are cases, in which the use of this article, under proper restrictions, might conduce to salutary results. It has always been regarded as a stimulating agent, and yet has been sometimes administered where stimuli were altogether improper.

The purest kind of oxygen should be used for inhalations, viz. that which is obtained from chlorate of potash; and bags of oiled silk or distended caoutchouc, furnished with ivory or silver tubes and stop-cocks, will present convenient vehicles for its administration. Some have given it in its pure, unadulterated form; and others have blended it with common air, and other gaseous matters. It has been tried in all stages of pulmonary consumption, and of course must have had very various effects. If useful at all, in those affections, it must be when inflammatory action has been removed, and great debility exists, combined with more or less difficulty of respiration.† It is reported to have done much good in asphyxia, induced by atmospherical impurities, and from the inhalation of positively noxious vapors.

We remark, in concluding our observations on oxygen, that the advocates of Sir Humphrey Davy's views, have attempted to

\* I am of the number of those who believe, that there is a constant chemical action occurring between the inspired oxygen and the carbon of the blood, throughout the arterial circulation. There is, probably, thus formed a gaseous oxide of carbon, which meeting with oxygen in the lungs, is changed into carbonic acid, and evolved in the act of expiration.

† For a particular account of its remedial effects, see Chaptal's Chemistry, page 61.



set aside the Lavoisierian system, partly, because discovery has shown that oxygen is not the sole cause of acidity, and that other bodies are alike capable of supporting combustion. This is a very unfair mode of inference, and not at all warranted by the case. There may have been a tendency to exclusiveness in the opinion of Lavoisier on this subject, because, in his day, oxygen was the only well known substance that appeared to be essential to acidity and combustion; but according to our views of the antiphlogistic chemistry, there is no feature of it at variance with subsequent discoveries, in relation to these points. The system is so perfect, in point of principle, as to admit of extension to meet all the probable advances of science.

---

## CHAPTER II.

### Of Hydrogen, and its Compounds with Oxygen.

THE word Hydrogen is derived from the Greek, and signifies, to *generate water*. I will not stop to dispute about the use of this term, now, although I am persuaded there are good objections against it, because it was doubtless, the best that presented itself when first adopted. The production of water always results from the mutual action of its elements, and both are alike concerned in the process.

This gas was formerly denominated *inflammable air*, because highly combustible. It was called *phlogiston*, from the supposition, that it was the matter of heat. As early as 1766, Mr. Cavendish made some researches into its nature and announced its distinctive properties. Its equivalent, by weight, is always regarded as 1, compared with oxygen as 8; by volume, it is one measure,  $\square$ . Its specific gravity is 0.0694; and as it is 16 times lighter than oxygen, it follows that 100 cubic inches weigh 2.118 grains. To arrive at this result, we have only to divide the weight of 100 cubic inches of oxygen gas, (33.888,) by 16.

Hydrogen gas refracts light powerfully, and is sparingly absorbed by water, which dissolves about a seventy-fifth part of its volume. This element is largely diffused over the surface of nature, constituting a ninth part of the waters of the globe, and being a component part of almost all the products of the animal and vegetable kingdom. When perfectly pure, it is a transparent, colorless, inodorous, insipid gas, distinguished by its unequalled levity and great inflammability.



It is impossible to prepare hydrogen, independently of the decomposition of water. The component parts of this fluid are oxygen and hydrogen; and in procuring the latter, the former is abstracted by some substance present, which has a powerful attraction for it. We may employ either iron or zinc for this end, though iron, in the state of filings or turnings, is usually preferred. We may put the metal into a glass retort, supported on a lamp stand in the way represented by Fig. 1st, Plate 2d, or into a common oil flask, to which a bent glass tube has been affixed in the mode described in the remarks explanatory of the plates. The extreme end of the tube is to be dipt under the mouth of an inverted glass receiver, for the purpose of collecting the gas. When the apparatus is ready, sulphuric acid, diluted with about five parts of water, is poured over the metal, and the gas very soon passes off.

The young operator should be cautious not to have more materials in the retort or flask, than will fill one third or half at most, of the body of the vessel, otherwise the rapid evolution of the gas may occasion the liquid to boil over, and so do considerable mischief. The greater the division of the metallic particles, and the larger the surface on which the diluted acid acts, the more freely will the gas be formed and disengaged. This teaches us, that the filings of iron are much more suitable than nails or other pieces of the metal, and illustrates a doctrine already inculcated, that the more effectually the attraction of cohesion is overcome, the more rapid will be the chemical action.

We have said that this gas is highly inflammable; and some one might venture to test the truth of this statement, by applying a lighted taper to the open end of the tube whence the gas is flowing, and without expecting it, might blow the flask and tube into a thousand fragments. Such accidents have happened, even in the hands of experienced operators, and considerable damage has resulted. Now, an explosion under these circumstances, is owing to precipitancy. Sufficient time should be allowed, to cause the hydrogen to drive out of the vessel all the common air, the oxygen of which, in union with the hydrogen, occasioned the explosion. A few minutes will suffice to accomplish the object proposed, and then the gas may be inflamed, with entire safety. A good mode for ascertaining whether hydrogen alone is escaping, is to fill a small jar with the gas over the pneumatic tub; then remove it on a plate or tray, keeping the mouth still downwards, and apply a lighted taper to it in that position. If atmospheric air be present, the contents of the jar will burn rapidly; but if hydrogen, only, it will burn very slowly, and the

combustion will be confined to the place where it comes in contact with the common air.

If zinc be employed, in place of iron, the finer the particles the better. The solid metal may be rasped, and the filings made use of. Some prefer what is called granulated zinc, which is made by melting the metal before a blowpipe, or in an iron ladle, allowing the fused zinc to fall into a vessel of cold water.

The most economical method for preparing hydrogen, is by transmitting the vapor of water over fine filings or turnings of iron. To this end we use a common brick or sheet iron furnace, having an aperture through its sides to admit the free passage of a gun-barrel, open at both ends. The filings must be placed in the barrel, so as to have them at the centre, as much as possible, but not so compact as to obstruct the free motion of the steam. To one end of the barrel, a bent tube of lead or other suitable material is attached, in order to convey the gas into receivers placed on the pneumatic tub or elsewhere. To the other end, some operators lute a tubulated retort, half full of water, which, on the due application of heat, furnishes the requisite steam. Instead of the retort, I make use of a tin cup, with a leaden tube passing from the bottom and accurately fitted to the gun-barrel. Just below the cup is a stop-cock, by a turn of which, a sufficient quantity of water can be let down to the filings. When the gun-barrel is at a red heat, the steam, sent over from the retort, or instantly formed from the water falling from the tin cup, is instantly decomposed; its oxygen combines with the iron filings, while its hydrogen escapes in the form of gas.

It is not necessary to expose the gun-barrel to a very high degree of temperature. A red heat is amply sufficient, as all kinds of iron tubes are soon destroyed in the open air at a higher temperature. As soon as the gas is collected and the bent tube removed, the gun-barrel should be taken out of the furnace and gradually cooled.

When this method of making hydrogen gas is adopted, for the purpose of filling the half of a compound blowpipe, or a larger receiver of the pneumatic tub, we determine the quality of the gas by applying a lighted taper to the open end of the gun-barrel, previously to attaching the bent tube. If too much water be let down from the tin cup, the excess of steam will extinguish the taper, but if pure gas alone be evolved, it will be kindled instantly, and burn with a long blaze.

A very convenient contrivance for preparing a small quantity of hydrogen, whenever it may be required, is represented in Fig. 9th, Plate 2d, having two apertures or necks, furnished with

ground glass stoppers. To one neck, we may fix, in lieu of the stopper, a glass tube of six or seven inches in length and one fourth of an inch in diameter. In the other aperture, we pass the zinc turnings and then add the diluted sulphuric acid. The gas soon escapes, and after waiting a few minutes to suffer all the common air to pass out, the steam issuing from the tube may be kindled. This constitutes the *philosophical candle*. The glass stopper should be fixed in its place, as soon as the diluted acid is poured in. If a lighted taper be applied to the gas as it issues from the tube, before all the common air is expelled, an explosion will occur, owing to a portion of oxygen that is present. As this might be a little hazardous, the precaution mentioned should always be observed.

The quantity of sulphuric acid necessary in the formation of hydrogen gas, depends upon the kind of metal employed. If we take zinc, 34 parts must be used for every 40 parts of real sulphuric acid, which are contained in 49 of the common liquid acid. One part of hydrogen gas will be evolved, from the 9 parts of water present, which consist of 8 oxygen and 1 hydrogen. Thus, before decomposition, we have 34 parts of zinc, 9 of water, and 40 of sulphuric acid, making in all 83 parts. The evolution of 1 part of hydrogen gas, reduces the amount total to 82, constituting the sulphate of zinc. The 8 parts of oxygen in the 9 parts of water, join the zinc, and form 42, oxyde of zinc, which uniting with the sulphuric acid, give 82 parts, as already stated, of sulphate of zinc.

In actual practice, however, we add a little more water than a strict conformity to the doctrine of equivalents would seem to require. Little or no action will take place on the zinc, if this precaution be not observed, and the sulphate of zinc that is formed, must have a sufficient quantity of water to retain it in solution, and so prevent it from falling in the form of powder and impeding the farther action of the metal on the liquid. As every 49 parts of common sulphuric acid contain 9 of water, we may use, by weight, 7 parts of zinc, 10 of sulphuric acid, and 60 of water. But if iron be used, the same quantity of sulphuric acid and water may be mixed with 6 parts of this metal, and the theory already stated, is applicable here. The decomposition of the water, furnishes oxygen to the iron, and an oxyde is the result, while the hydrogen escapes. The acid joins this oxyde, and a sulphate of iron is formed. Every 28 parts, (one equivalent,) of iron, act on 9 parts of water and 40 of sulphuric acid, just as 34 parts of zinc did in the other case.

In the mode of making hydrogen gas, by the transmission of watery vapor over red hot iron filings, every 28 parts of iron,

(one equivalent,) decompose 9 parts of water, forming 36 of oxyde of iron, which remain in the tube, while one part of hydrogen is disengaged.

If it is desirable to determine, how much hydrogen gas is formed by the action of a given weight of zinc or iron, one grain of gas may be allowed for every 34 grains of zinc, or 28 of iron, employed in the process. As the weight of 100 cubic inches of this gas is 2.118 grains, so one grain will be the weight of 47 cubic inches, or very nearly. Then, if we multiply the weight, in grains, of the hydrogen that ought to be disengaged, by 47, we shall find the exact volume, in cubic inches, which it will occupy in the gaseous state.

It has been remarked, that perfectly pure hydrogen gas is inodorous, insipid, &c., but it is equally true, that we seldom meet with it free of a disagreeable odor. It often contains some admixture of sulphuretted hydrogen, and as zinc is sometimes contaminated with arsenic, there is occasionally formed, a portion of arseniuretted hydrogen. So also, some traces of carburetted hydrogen are now and then discernible; and from these sources, it is easy to perceive that not a little impurity must be apparent. Most of the iron used in forming hydrogen, contains more or less carbon, and the filings are often blended with oily particles, which contain carbon, also, and hence arises the formation of carburetted hydrogen gas. The zinc, in addition to small portions of arsenic, also contains a little sulphur, and this furnishes the sulphuretted hydrogen. If the zinc be distilled previously to the preparation of hydrogen gas, it will be freed from some of the impurities common to it, and then it will yield a gas nearly pure.

In addition to the apparatus for making hydrogen gas already mentioned, another deserves to be noticed, called the *self-generating hydrogen machine*. I am in the habit of using it in my lectures, but as it cannot be explained satisfactorily, without a plate, and as the other modes for preparing this gas answer sufficiently well, it is needless to dwell on it in this place. Its chief advantages are the saving of zinc and acid, and the readiness with which the gas may be formed, at any moment.

Hydrogen gas burns with a pale, watery looking flame, which affords a very feeble light. To become acquainted with the peculiar appearance of the flame, apply a lighted taper

Exp. to it as it issues from a gas holder, or a bladder with a long tube attached to it. The delicate blue flame will thus be very obvious.

Fasten a piece of candle to an iron wire and introduce it into a jar of hydrogen gas, having the mouth downwards; the gas will be inflamed and continue to burn at the mouth of the jar,



while the candle will be extinguished so soon as it is immersed in the gas. This shows conclusively, that hydrogen, Exp. although highly inflammable, cannot burn, apart from oxygen. At the surface of the hydrogen, atmospheric air is present and furnishes the requisite oxygen, and hence the combustion in that particular spot.

If we fill another jar and place the mouth upwards, and then apply a lighted taper, all the gas will be consumed speedily, because it rises by its natural levity, and mingles with the oxygen of the air. If we wait a very short time, before Exp. applying the taper, we shall discover that all the gas has escaped and of course there will be no combustion.

In these experiments, and in all others where hydrogen is burned, it unites, in the act of combustion, with the oxygen of the air; water is formed, and the nitrogen previously combined with the oxygen in the atmospheric air, is disengaged.

If we mix, either in a bottle or bladder, one measure of hydrogen with half its volume of oxygen gas, and fire the mixture by means of a lighted taper, a loud explosion will ensue, and watery vapor will be formed, resulting from the union of the two gases. If a bottle be used, it should be a strong one; a black porter bottle will answer, and it should be wrapped round with a thick towel, to save the operator and spectators from injury. The reason for mixing the gases in the proportions above named, will appear by reference to the table of equivalents, by weight and measure.

	Hydrogen.	Oxygen.	
Equivalents by weight,	1	+	8 = 9 water.
Do. by volume,	□	+	□ = □ watery vapor.

Thus it is plain that one equivalent of hydrogen, by weight, corresponds with one measure or volume of this gas, and that one equivalent, by weight, of oxygen, 8, agrees with a half measure or volume; this gas being 16 times heavier than hydrogen. The explosion of one volume of hydrogen and a half volume of oxygen, mixed, gives rise to one volume, and no more, of water or watery vapor; whence it results, that a half volume has been lost by condensation. To prove that these gases combine exactly as we have stated, we need only say, that if either be added in excess, the surplus gas uniformly remains, in its true character, after the detonation.

Mixtures of hydrogen with atmospheric air also explode, as may be shown with the air pistol. If this be partly filled Exp. with hydrogen gas, the atmospheric air in the pistol will afford enough oxygen to cause a detonation. But as at-



mospheric air contains only a fifth part of oxygen gas, every measure of hydrogen gas will require two and a half measures of air for its complete combustion.

For the purpose of firing hydrogen, in these experiments, some persons prefer a strip of paper rolled up into the form of a match; others use an iron wire heated to redness; and many employ the common wax tapers. I have used them all, and employ the one or the other, just as it suits my convenience.

Procure a glass tube open at both ends, two feet in length and tapering in width from a half inch to two inches, and place it over the flame of the *philosophical candle*, already noticed, and  
 Exp. a pleasing musical sound will be produced, admitting of considerable variations. This phenomenon is supposed to result from the feeble and quickly repeated explosions of the inflammable gas.

Over the same gas, in a state of combustion, place another tube of similar size, but closed at the upper end, and as the gas burns, moisture will be deposited on the sides of the tube, owing  
 Exp. to the formation of water, by the union of the inflammable gas with the oxygen of the atmospheric air.

A reduced pressure is unfavorable to the combination of oxygen and hydrogen gases, while an increase of the same power causes them to unite more readily. If the mixture be made in due proportions, and sudden pressure be applied, combination instantly results, attended with a loud explosion; and when the operation is performed in a vessel of glass, a bright light is perceptible and the vessel is fractured. Biot was the first to make this experiment, which is necessarily attended with hazard, and is, withal, a difficult operation. As a very considerable amount of heat is evolved in the act of compression, it is quite probable, that this agent is the efficient cause of the chemical union effected between the gases.

It may be stated here, that a few drops of sulphuric ether, added to a quantity of oxygen, will give rise to a gaseous mixture, which may be inflamed by a red hot iron wire. In like manner, a mixture of oxygen and hydrogen may be fired by similar means, or by using a red hot coal, and the explosion  
 Exp. will be very perceptible. If we employ a coal, or iron wire, at a *dull* red heat, the gases will combine, but there will be no detonation.

What is denominated *spongy platina* has the power of inflaming hydrogen gas. (This article is the metallic platina reduced to a state of minute division, as will be more fully explained hereafter.) If we force a stream of hydrogen gas, either from a bladder furnished with a stop-cock and tube, or from

the self-generating hydrogen machine, or in any other way, directly upon a piece of the spongy metal, in contact with oxygen gas, or atmospheric air, it will be instantly ignited and the hydrogen itself inflamed. This result is the more certain, if the platina be perfectly dry; should it be damp, it will be proper to heat it in a crucible, or to expose it to the action of a spirit lamp, on a metallic plate.

Several contrivances have been made, based on this property of spongy platina, in relation to hydrogen gas, for preparing night lamps, and for other purposes, requiring instantaneous light. In all of them, the materials for forming hydrogen gas are present, and the gas, when formed, is allowed to escape suddenly from a very small orifice, and to fall on the spongy platina.

We have alluded to the application of this principle, by means of a bladder furnished with a stop-cock and pipe, and the effect may thus be shown very satisfactorily. A pneumatic jar, also provided with a stop-cock, must be at hand, which can readily be screwed to the stop-cock of the bladder. (See Figs. 13th and 14th, Plate 1st.) The two instruments being connected, and the stop-cocks shut, the jar is to be filled with water and inverted on the shelf of the pneumatic cistern. It is then to be filled with hydrogen gas; after which, by opening both stop-cocks and depressing the jar into the water, the bladder is quickly filled. We then shut the stop-cocks, detach the bladder from the jar, and having affixed the pipe, the gas may be forced against the spongy platina, and exhibit the phenomena already stated. Exp.

But the spongy platina not only inflames hydrogen gas; it acts with equal power on the mixture of this gas with oxygen. When these gases are in the due proportions for forming water, a small piece of spongy platina will as certainly inflame the mixture, as would a lighted taper, or an iron wire at a bright red heat. This experiment may be conducted with entire safety, by using a bladder for the mixed gases. The spongy platina causes an instantaneous explosion, and as the bladder is blown to pieces with much violence, it is well to have the hand, that is most exposed, protected by means of a glove.

Mr Reid advises to fit a large cork to the neck of the bladder, having previously made a hole in the centre; an object which is best attained by passing a red hot iron cylinder through the cork. This hole is designed to receive another cork, of small size, to which an iron wire is fastened, having at its lower extremity a small contrivance for containing the spongy platina. The bladder may be tightly secured around the large cork, which will thus constitute a kind of neck, and it Exp.

will be no difficult matter to fill it with the mixed gases, and to close the aperture. Let it then be removed to a proper stand, and having the small cork with its iron wire and platina appendages at hand, let this be quickly introduced, so as to avoid the escape of any considerable portion of the contents of the bladder. In a very short time explosion will follow.

If other gases be present, in the mixture of oxygen and hydrogen, detonation will either not take place at all, or some considerable time will elapse before its occurrence. But the spongy platina does not fail, even here, to effect a combination, and a sufficient degree of heat is evolved to ignite the metal. All these experiments with this article, have been performed, occasionally, with a mixture of the spongy platina and clay, made into small balls.

As the combining proportions of oxygen and hydrogen, in order to form water, are, a half measure of the former to a whole measure of the latter, it is plain that the proportion of either oxygen or hydrogen, in mixed gases, may be determined by adding one or other of these to the mixture, and observing the extent of condensation that results from their union. Suppose it is desirable to know the degree of purity of a quantity of hydrogen gas, made according to any of the processes already laid down. We mix a given portion of the gas to be examined, with an equal quantity of oxygen gas, in a glass jar over mercury, and then introduce a mixed ball of spongy platina and clay, so as to inflame without an explosion. Condensation succeeds to the union thus effected, and if it be equal to the whole bulk of the hydrogen and one half of the oxygen employed, it follows that the hydrogen was pure, because these are the precise proportions in which these gases combine to form water; if a less degree of condensation be apparent, the extent is to be marked down accurately, and two-thirds of the amount will give the precise quantity of pure hydrogen contained in the gas under examination.

Just in the same way, is the purity of oxygen gas determined, with sufficient accuracy. We mix it with an equal bulk of pure hydrogen gas, and effect their combination as before, noting accurately the degree of condensation. If the loss of bulk in the mixed gases shall prove to be equal to three times the volume of the oxygen employed, it follows that this gas must have been pure, since it naturally combines with double its bulk of hydrogen. But should the loss in bulk be less, we take one third of its actual amount, as indicating the precise quantity of pure oxygen that was present.

By passing a succession of electrical shocks, or a stream of

galvanism, through a mixture of oxygen and hydrogen gases, they may be inflamed, and water will be the result. The instrument proper for this experiment may be seen in Plate 2d, Fig. 7th.

The combustion of hydrogen gas alone, gives rise to a very strong heat, but if it be mixed with oxygen, in the proportions for forming water, and the mixed gases be burnt as they issue from a small orifice, we then have a more powerful heat than can be obtained in any other way, excepting, perhaps, by the galvanic machines of greatest intensity. It is on this principle that the compound (or oxy-hydrogen) blowpipe is constructed; an instrument, for which we are indebted to the ingenuity of Professor Hare, of the University of Pennsylvania.\* It consists of two gas holders, made of sheet copper, and joined together by means of copper hoops. At the distance of ten or twelve inches above these vessels, is a copper basin large enough to contain several gallons of water, which, by turning the stop-cocks, may be made to pass through cylindrical tubes, down to the gas holders. These latter are first filled with water by an aperture in each, within a half inch of the bottom, and as atmospheric pressure is removed from above by the closure of the stop-cocks, the water will not escape after the vessels are filled, until driven out by means of the gases. Each apartment is filled with its appropriate gas, hydrogen on the one side and oxygen on the other, and then the lower apertures are closed with good corks. Within an inch of the top of each gas holder, a stop-cock is fixed, and to these an instrument is attached, having two hollow tubes joined at the smaller end, so that both terminate by one very small aperture; and as the heat at this point is very intense, when the gases are burned, a small piece of platina or gold is usually placed there, in preference to a metal of more easy fusion. When the compound blowpipe is about to be used, the stop-cocks are turned by which the water from the basin is to descend into the gas holders. This will necessarily give impetus to the gas whenever it can find vent. At this moment the hydrogen stop-cock is turned, and the gas inflamed as it rushes out, and then the stop-cock of the other side is opened, which gives passage to the oxygen and so augments the intensity of the heat.†

\* The attempts of some British chemists to defraud the real inventor of this discovery, have been so frequently exposed, that we deem it a waste of time to dwell on the subject here.

† A blowpipe of considerable power may be had, by using the oxygen side only.



This operation may be performed without hazard, because the gases are not only escaping with force, but they are fired at a distance of several inches from their respective chambers; and so long as the water presses with sufficient power to drive them out, the flame cannot pass within the small orifice where the gases are first kindled. In Clarke and Newman's blowpipe the oxygen and hydrogen gases are contained in a common vessel. The mixed gas is compressed to such an extent, that on opening the stop-cock it rushes out with violence, and with all the care that is employed, the machine is sometimes blown to pieces. If a bladder furnished with a stop-cock and small tube, be filled

Exp. with the mixed gas and a board be laid upon it, with suitable weights, so as to compress the bladder as much as possible without bursting it, and the gas be inflamed as it issues from the fine point of the tube, it will burn without an explosion. But if there be not a sufficient weight on the board to insure the requisite impetus to the gas, there will necessarily be a detonation.

In the use of Professor Hare's blowpipe, I have never met with an accident, and regard it as one of the safest implements in the laboratory. The heat generated by it, arising from the

Exp. combustion of oxygen and hydrogen gases, is sufficient to melt all the metals, barytes, strontian, lime, potash, and many other substances which were formerly called infusible.

The use of hydrogen in filling balloons, arises from its great levity. This property is easily shown, by what are called hydrogen soap-bubbles. The experiment is not a mere plaything, but actually illustrative of characteristic properties. Fill the stop-cock bladder (Fig. 13th, Plate 1st) with hydrogen, and fasten a tobacco pipe to the mouth of the stop-cock. Dip the bowl of the pipe into a lather of soap, then press the bladder, and bubbles will rise in rapid succession. As these are floating in the air, touch them with a lighted taper and they will burn with a feeble report. If the bladder contain oxygen and hydrogen gases, the bubbles will not rise so easily, (because the mixed gas is heavier than hydrogen alone,) but they will explode with a loud noise.

Some experiments, performed by Chaptal, would seem to show that hydrogen gas may be respired for a few moments, and not prove fatal. If, however, an animal be confined in an atmosphere of this gas, death will result in a very short time. It is fair, therefore, to class it with the irrespirable gases. In the days of pneumatic medicine, hydrogen was resorted to, in a mixed form, and its action was judged to be sedative. It soon, however, was laid aside, and there is little reason to believe that



it will ever be restored. An individual, by name Maunoir, who succeeded in inhaling hydrogen for some minutes, experienced a peculiar effect from it on his voice. It became soft, shrill, and even squeaking, to such a degree as to alarm him, and that, too, without in any perceptible degree impairing the general powers of his system. See Tilloch's Magazine, Vol. IV.

There are two compounds resulting from the union of oxygen and hydrogen. The first is the natural or native production, called *water*; and the second, which is always artificial, is denominated the *deutoxyde of hydrogen*, and by some, *oxygenated water*.

### *Of Water.*

The equivalent number for water is fixed at 9, because it is a compound of 1 hydrogen and 8 oxygen, by weight. Its equivalent, by volume, is *one* measure, because in forming it from a *half* measure of oxygen and a *whole* measure of hydrogen gases, there is a condensation of a half measure in the chemical action that ensues. It is made the standard for estimating the specific gravity of bodies, and is called 1 or 1000. A cubic inch of water weighs nearly 253 grains, is 828 times heavier than atmospheric air, and when changed to vapor by a heat of  $212^{\circ}$ , it is expanded to about 1700 times its ordinary bulk. The specific gravity of its vapor, (or steam,) is 0.625.

We have spoken already of the formation of water by the detonation of oxygen and hydrogen gases, and also by the burning of hydrogen gas within a glass tube. We have also seen that water is decomposed into those very gases, in the formation of hydrogen, by passing steam over red hot iron filings. In like manner, this fluid may be resolved into its constituent gases by the agency of galvanism. Thus, if two glass jars be filled with water and inverted over that fluid in a glass basin, (in order to see the experiment the better,) close to each other, and if the positive wire from a galvanic battery be passed under one jar, and the negative wire under the other, it will be seen, that gas will rise into each jar and displace the water. On examination, it will appear that one of the jars contained hydrogen and the other oxygen. We are thus apprised, by analytical as well as synthetical proof, of the real composition of water.

When the gases composing water are combined at the temperature of  $212^{\circ}$ , and the same heat is continued, we shall find that condensation has taken place, just equal to the quantity of oxygen employed; and of course that the volume of the steam generated is just equal to the original bulk of the hydrogen, at

212°. In every point of view in which water is contemplated, it must be regarded as an agent of great importance. It pervades very many substances, in which its liquid character is not visible, and these are usually called *hydrates*; it also combines with a large number of articles, which it reduces to its own fluid form, and hence its uses as a solvent. Its presence is often indispensable to chemical action; and there are a few operations in which its absence is essential to entire success.

The experimentalist requires pure water, because the foreign matters frequently blended with that fluid could not fail to affect his results. It is proper, therefore, not only to state some modes by which pure water may be obtained, but also to designate the characteristics of that fluid, in a state of purity. It follows, as a matter of course, that water must partake more or less of the nature of the soil or strata, through which it percolates; and it may be affirmed with certainty that a water has this or that peculiar character, if we know precisely through what sort of matter it has travelled.

Melted snow, or rain, collected at a considerable distance from populous places, and where the atmosphere is unvitiated by offensive vapors, will give us the purest water that can be had, apart from artificial operations. It will be found, however, that even this kind of water contains common air, and fixed air, (carbonic acid gas,) which may be disengaged, either by boiling, or placing it under the exhausted receiver of an air-pump.

The chemist does not depend on such sources as these when he requires pure water, but resorts either to repeated filtrations or to distillation, the latter of which, only, can furnish him with the article he needs. Some persons use the common still for this purpose, and it will answer very well; but I find it more convenient to get pure water by means of a glass retort and receiver. (See Fig. 8th, Plate 2d.) The water is placed in the retort and heated by means of a chauffer, spirit lamp, or sand bath. The vapor, as it passes over into the receiver, is condensed by the application of cold cloths. Before the retort is luted to the receiver, the water should be boiled for two or three minutes, in order to expel any gaseous matter that might be present, as well as to remove particles attached to the sides of the retort. It will be perceived that this process is, essentially, a distillation.

The water thus obtained is transparent, colorless, and very limpid; it is insipid, and therefore unpleasant. When poured from one vessel to another, it has none of that sparkling appearance which fresh well-water exhibits, because it is deprived of

air. Its specific gravity is less than that of the water in the source, when it was originally taken, because it is freed of all saline and other particles which naturally increase the density of that fluid. It is not *hard*, as before distillation, for the same reason, and it mixes easily with soap, giving no precipitate when added to solutions which indicate the presence of foreign matters. Another property said to belong to pure water is, that it moistens substances more readily than undistilled water. Mr Perkins has shown, by decisive experiments, that water is, to a certain extent, compressible. It is probable, however, that this property is not so great as his researches seem to indicate.

It must be obvious, that in order to keep distilled water free from gaseous or other admixture, it should be kept in perfectly tight vessels, and this end is attained by making use of clean glass vessels with ground stoppers that fit accurately. Dr Henry has made out the following table to show how large a quantity of several different gases can be absorbed by 100 cubic inches of water, at the temperature of 60°, and 30° of the barometer.

	According to Dalton.		According to Saussure.
Sulphuretted hydrogen,	100	cubic inches	253
Carbonic acid,	100	do.	106
Nitrous oxyde,	100	do.	76
Olefiant gas,	12.5	do.	12.3
Oxygen gas,	3.7	do.	6.5
Carbonic oxyde,	1.56	do.	6.2
Nitrogen gas,	1.56	do.	4.1
Hydrogen gas,	1.56	do.	4.6

There are many purposes for which water passed through a paper filter will answer sufficiently well. The common class experiments may be performed with water thus prepared, when much nicety is not required in the results. But in all delicate processes, where great accuracy is desired, it is indispensably necessary to use water of the purest kind.

It is not only in the circumstances already detailed, that water undergoes decomposition. The compounds of sulphur and metals, usually called sulphurets, readily decompose water. The sulphur and metal have a disposition to unite with the two elements of that fluid; the sulphur joins the hydrogen and a gas, called sulphuretted hydrogen, is the result, while the oxygen and metal combine to form an oxyde. The other simple substances, which, in their union with metals, give rise to bodies analogous to oxydes, as for example, chlorides, iodides, bromides,

cyanides, &c., all have a strong affinity for hydrogen. Hence, the chlorine in a chloride will separate hydrogen from water, and form an acid substance, while the metal, taking the oxygen thus liberated, is changed into an oxyde, and this enters into combination with the newly formed acid.

We have several compounds of hydracids, (that is, acids which contain hydrogen, but no oxygen,) and metallic oxydes, which, by decomposition, give rise to a portion of water; in these cases the hydrogen of the acid unites with the oxygen of the oxyde, and the product follows.

When treating of caloric, we noticed its action on water, and the large amount of latent heat necessary to the existence of steam. Dr Ure has made an estimate on this subject, from which it appears that the quantity of caloric actually absorbed by water in its passage to the state of vapor, would raise the temperature of that liquid nearly one thousand degrees, if it could remain in the fluid state. Now all this latent heat is evolved and of course becomes sensible to the thermometer, when the steam is condensed, as in passing it into a vessel of water; in this way a large bulk of water may be speedily warmed.

For drying precipitates by steam, and for other purposes, where the direct application of fire would be hazardous, or otherwise improper, I use a simple apparatus, invented by Dr Ure. It consists of a tin tea-kettle, to the mouth of which a box is fastened by means of a tube. This box is also of tin, about 12 inches long, 12 wide, and 6 deep. In the top are apertures for funnels containing precipitates, and also for flasks, &c. When the water in the kettle is boiled, the box is filled with steam, which acts on the vessels placed in the top of the box. It is precisely on this principle that steam is applied to the drying of gunpowder. Water in the solid state, in which it is found in nature, is the well known substance, ice. This occupies a larger bulk than the water did previous to the act of freezing, and is regarded as an exception to the law, which recognises a contraction in the passage from the fluid to the solid state. But in the case of ice, it must be remembered that there is a crystalline arrangement of such a nature as necessarily to augment the bulk, and with this we also find the existence of cells which contain air, a fact that serves to explain in part the levity of ice. The presence of air in this substance, may easily be shown by sinking a piece to the bottom of a glass tumbler nearly full of water, and confining it there. What is called *rotten ice*, (on the breaking up of winter,) sinks, because these cells are obliterated and the air discharged.



The subject of *mineral waters* will be noticed at some length in another part of the work. As they contain substances which have not been considered, nor even mentioned, it is deemed advisable to devote a short section to their character, properties, and analysis, after the student shall have learned something of the various saline and other bodies that enter their composition.

The second and last compound of hydrogen and oxygen, or that which may be denominated *artificial*, is the DEUTOXYDE OF HYDROGEN, OR OXYGENATED WATER.

It contains, for every equivalent of hydrogen, or 1, two equivalents, or 16, of oxygen, and hence it has been called the *deutoxyde*.\* Its equivalent is, of course, 17, and its specific gravity 1.452.

This substance may be prepared by the following process, viz. dilute muriatic acid with ten times its bulk of water in a glass vessel, surrounded with ice or snow, and add finely powdered deutoxyde of barium, (to be more particularly noticed hereafter,) in small portions at a time, as long as the acid will dissolve it. Then sulphuric acid, (common oil of vitriol,) is poured into the solution to effect the decomposition of the muriate thus formed. This acid, in combining with the barium, separates one portion of oxygen from the deutoxyde, so that there is a sulphate of the protoxyde formed. As often as this deutoxyde is added and treated in the way already mentioned, one equivalent of oxygen is evolved from every equivalent of deutoxyde, and the water becomes in this way oxygenated. This process is repeated until the water takes up from twenty to thirty times its bulk of oxygen gas.

The deutoxyde of hydrogen, thus prepared, is still diluted with a considerable quantity of water, which must be removed by placing the vessel containing it under the exhausted receiver of an airpump, along with another vessel of strong sulphuric acid. The action of the pump evaporates the water, which is instantly taken up by the sulphuric acid. The evaporation must not be carried too far, or the deutoxyde will also be evaporated and mixed with the water and sulphuric acid.

As yet, we know of no useful purpose to which this deutoxyde has been applied. It is distinguished by the great facility with which it is decomposed by most of the metals and some of the metallic oxydes, especially the oxydes of silver, lead, gold, mercury, and platina. The decomposition ensues directly after

\* Some writers call it the *peroxyde*, because it is the highest state of oxydation of which hydrogen is susceptible.



they are brought into contact, great heat is evolved, and the oxydes are reduced to the metallic state, if they belong to the class of metals whose oxydes are reducible without the aid of carbonaceous matter. The deutoxyde of hydrogen is a colorless, transparent liquid, without odor. It whitens the skin, occasions a prickling sensation, and destroys its texture if kept in contact for a long while. It has a similar effect on the tongue, thickens the saliva, and gives a taste like some metallic solutions. It destroys the color of litmus and turmeric papers, and makes them white; of course, it may be said to possess bleaching properties.

A temperature of  $60^{\circ}$ , Fahrenheit, decomposes this compound, resolving it into oxygen gas and water; and if it be suddenly exposed to a heat of  $212^{\circ}$ , the oxygen is disengaged with explosive violence. When it is united to water or the stronger acids, it retains its oxygen more firmly.

---

### CHAPTER III.

#### Of Nitrogen, and its Compounds with Oxygen.

THE name *nitrogen* has been given from the circumstance of its being essential to the formation of nitric acid. Its former name, *azote*, was applied by Lavoisier, on account of its fatality to animal life; but as other gases are equally destructive of vitality, there seemed to be a want of definiteness in the term. Dr Rutherford appears to have been the first to notice this substance; his experiments were made in 1772. Not long after, it was investigated by Lavoisier and Scheele, both of whom concluded, that it was an essential constituent of the atmosphere.

The combining number or equivalent of nitrogen appears to be 14, by weight, and *one* measure, by volume. Its specific gravity is 0.9722. One hundred cubic inches weigh nearly 30 grains. It is scarcely absorbed by water; certainly in less quantity than oxygen.

In its original state, nitrogen is gaseous, but loses this form, in some instances, when it enters into combination. It has no active, positive properties, by which it can be distinguished from other gases. It is transparent, insipid, inodorous, colorless, incombustible, irrespirable, and incapable of maintaining combustion. It constitutes, all over the world, about four-fifths of atmospheric air; it is a component part of almost all animal

matters, and is sometimes found in vegetable bodies. The article with which it is united in the air we breathe, has already been spoken of, viz. oxygen, though a small fraction of another gaseous substance is also present, which will be noticed in due time. Besides the air of the atmosphere, which some regard only as a mixture, nitrogen enters into several acids and oxydes. Thus we have the nitrous and nitric oxydes, which are gaseous, and the nitric, nitrous, and hyponitrous acids. With hydrogen we shall see that nitrogen forms ammonia; with carbon, cyanogen; besides which there are several interesting compounds, of which nitrogen makes a part.

Any process, by which we can abstract all the oxygen from a given bulk of common air, will yield us tolerably pure nitrogen gas. One of the most common methods is, to burn the substance called phosphorus, in atmospheric air, in a jar or bottle, over water. Having the water a half inch deep on the shelf of the pneumatic cistern, place on its surface a small piece of wood with a very little tin or copper cup on it, so that the whole will readily float. The cup need not be larger than the bulk of a few grains,\* and is intended merely for a receptacle of the phosphorus. Having an inverted jar at hand, containing nothing but common air, kindle quickly the phosphorus in the cup, and instantly cover it with the jar. A large quantity of white fumes will appear, immediately, owing to the combination of the phosphorus with the oxygen of the air. These fumes are phosphoric acid, and for a time they are blended with the nitrogen of the air; but as water rapidly absorbs them, they soon disappear and the water rises within the jar, leaving the nitrogen tolerably pure. If the jar has a stop-cock attached to it, the gas can easily be transferred into another inverted jar filled with water, by simply depressing the former into the cistern; and when the stop-cock is brought under the other jar and turned, all the gas will rise. The gas in thus passing through water two or three times, is freed entirely of phosphoric acid. It contains, still, a small portion of carbonic acid, which may be removed by agitation with caustic potash. For all ordinary purposes, the gas may be used in a few hours after the combustion of the phosphorus, without purification.

In thus forming nitrogen, some persons make use of a tin stand with a cup placed on it, but the floating wood is a better contrivance, because it rises precisely as the water is elevated within the jar.

The sulphuretted hydro-sulphuret of potash or lime, when

\* Five or six grains of phosphorus will answer for a quart jar.

dissolved in water, is also capable of abstracting oxygen from atmospheric air. The solution is poured into a bottle, so as to fill one third of it, and the stopper, having been greased with a little tallow, to prevent the escape of air, is accurately fitted in. The bottle is then to be frequently shaken, and in a few hours, the separation will be complete. To get the nitrogen, we take the stopper out of the bottle under water, and easily transfer the residual gas to an inverted pneumatic jar.

Atmospheric air may, in like manner, be divested of its oxygen, by the action of a mixture of sulphur and iron filings. If one part of the former and two of the latter be made into a paste with water, and exposed to a gentle heat before the fire, for two or three minutes, or rather until the mixture is completely heated and its color changed nearly to black, a sulphuret of iron is formed, which rapidly attracts oxygen. Put a portion of this compound into a tin cup placed on a piece of wood, and invert over it, and on the pneumatic shelf, a jar containing common air. The sulphur and iron are both operative in the work of abstracting oxygen from the air, and if the process were continued long enough, fresh air being supplied, and the sulphuret moistened, it would be at length changed into a salt, commonly called copperas, but more correctly, sulphate of iron. The oxygen absorbed, changes the sulphur into an acid, viz. the sulphuric; it also converts the iron into an oxyde. This sulphuric acid and oxyde of iron readily unite, and the salt, already named, is the result. But more of this hereafter.

We have said that nitrogen constitutes four parts in every five of atmospheric air; of course, the processes above detailed for its formation, will necessarily yield that proportion, when they succeed in abstracting all the oxygen. This results as a natural consequence.

Another method must be brought to view here, which is occasionally resorted to in the preparation of nitrogen. It is based, most probably, on the well known circumstance, that nitrogen is a component part of all animal matter. Muscular flesh is exposed to the action of nitric acid, diluted with three times its bulk of water, and, if necessary, a gentle heat may be employed. The gas is thus evolved in considerable quantities, and may be collected in a jar, inverted in the usual way. In this process there is always some carbonic acid formed, but it is too small in quantity to interfere with the use of the nitrogen, for ordinary purposes. The theory of the production of the gas, in this instance, is a matter of dispute. Some have conjectured, that the nitric acid is decomposed and the nitrogen evolved from that source. But as the remaining liquid will

neutralise as much alkali as the original acid would have done, it is perhaps a nearer approximation to truth to say, that the nitric acid simply destroys the affinities existing in the flesh, and so separates the nitrogen from the other component parts of the animal substance.

The following mode of making nitrogen gas, is from Silliman's Journal, vol. xviii. p. 259.

When the metal called zinc, is dipped into fused nitrate of ammonia, the former is instantly oxydised, by the oxygen of the nitric acid contained in the salt just named; the nitrogen of the acid is liberated, together with the ammonia, both in form of gas. The zinc disappears as rapidly as when acted on by the strongest mineral acids, and the heat thus excited, is more than sufficient to complete the process. A temperature of from  $280^{\circ}$  to  $300^{\circ}$  will answer the purpose, but a small portion of zinc soon raises it to  $540^{\circ}$ . We are assured by Professor Emmett, the author of this process, that neither nitrous oxyde nor nitrous gas are evolved; but simply nitrogen, mingled with ammoniacal gas, the latter of which is instantly absorbed by the water, through which the gaseous mixture passes in order to be collected in suitable jars or receivers. A tubulated retort is to be partly filled with the nitrate of ammonia, and a cork is to be fitted to the tubulure. Through this cork a fine knitting needle or iron wire is to be passed, terminating in a hook, so as readily to hold a small coil of zinc. A lamp heat soon fuses the salt, and then the needle or wire is to be pushed down, so as to bring the zinc in contact with the melted salt. This is a convenient, and, at the same time, a necessary arrangement; for if the zinc be forced into the fused salt, the action would be so violent as to be unmanageable. The metal must, nevertheless, be constantly kept in actual contact, in order to keep up the action; otherwise a vacuum might occur and the retort be burst, by the sudden absorption of water from the pneumatic tub. It is plain, that this process offers a method of getting nitrogen as pure, and perhaps more so, than any other.

I have thus brought together a variety of modes by which nitrogen is prepared, because it is desirable that the student should be fully informed on this department of our science. It will be found, however, more convenient in practice, to obtain this gas by means of phosphorus, as fully detailed above.

A very simple experiment will show that nitrogen gas is not a supporter of combustion. As it is very little lighter than common air, we may place a jar full of it with the mouth upwards, and dip a lighted taper into it. The flame is instantly put out.










Exp.

Exp. If the jar be allowed to remain uncovered for the space of five minutes, or even less, it will then be seen that a lighted taper will burn in it, as freely as in the open air.

Exp. The comparative effects of nitrogen and oxygen, may be shown by placing two jars containing these gases on a table, and dipping a lighted taper into them, in quick succession. The nitrogen, in most instances, totally extinguishes the light, but if a spark remain, when the taper is withdrawn, the flame will be rekindled by the oxygen gas. And if the operator be sufficiently expert, these opposite effects may be several times repeated with the same quantities of gases.

Exp. We form a gas which has all the properties of atmospheric air, by mixing four parts of nitrogen with one of oxygen; excepting, however, the small portion of carbonic acid, always present. This experiment confirms the account previously given of the unvarying constitution of the air that surrounds us.

The following are the compounds which nitrogen forms with oxygen, and also with hydrogen, and which are presently to be examined. These compounds are stated in connexion with the equivalents, by weight and volume, of the substances that enter into their formation.

	Nitrogen.	Oxygen.		Nit.	Oxy.
Atmospheric air,	28	+ 8 = 36	or		+  = 
Protoxyde of nitrogen,	14	+ 8 = 22	or		+  = 
Deutoxyde of do.	14	+ 16 = 30	or		+  = 
Hypo-nitrous acid,	14	+ 24 = 38			
Nitrous acid,	14	+ 32 = 46			
Nitric acid,	14	+ 40 = 54			
	Nitrogen.	Hydrogen.			
Ammonia,	14	+ 3 = 17			

#### SECTION I.—PROTOXYDE OF NITROGEN.

This gas has had several appellations, but the one now in general use among accurate chemists, accords best with the correct principles of chemical nomenclature. Its discoverer, Dr Priestley, named it *dephlogisticated nitrous air*. The more familiar name of *nitrous oxyde* was given by Sir Humphrey Davy.

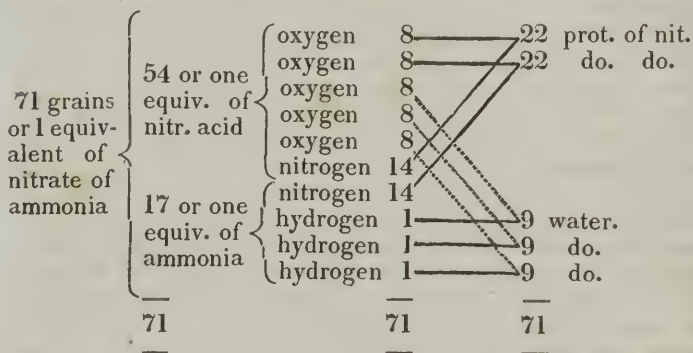
It will be seen, that in all its forms of combination, excepting in atmospheric air, nitrogen observes a uniformity of proportion. Its equivalent or combining number is always the same, while the proportion of oxygen varies in each compound. Thus, in the protoxyde or first degree of the oxydation of nitrogen, there



are 14 parts of nitrogen and 8 of oxygen, making the equivalent of the protoxyde 22, by weight. By volume, this compound is *one* measure, although made up of a half measure of oxygen and a whole measure of nitrogen; there is, of course, a condensation of a half measure. In the subsequent compounds, it is seen that the nitrogen continues at 14, while the oxygen is doubled, trebled, quadrupled, &c. exactly, without a fraction; and this is what is meant by *multiples by some whole number*. The specific gravity of the protoxyde of nitrogen is 1.527. It is made up of one volume of nitrogen, whose specific gravity is 0.9722, and a half volume of oxygen, which gives 0.5555 for its specific gravity, and the two added furnish, precisely, the specific gravity of the protoxyde. One hundred cubic inches weigh nearly 47 grains, and water absorbs more than half its bulk of this gas. The protoxyde may be formed, by exposing the deutoxyde or nitric oxyde for some days to the action of iron filings, or other substances, which, having a strong affinity for oxygen, take one equivalent of this substance from the deutoxyde, and thus reduce it to the state of protoxyde. This is not, however, the ordinary method for preparing this gas. It is usually obtained by decomposing the salt called nitrate of ammonia. This result is had by exposing the salt to heat in a glass retort, water being formed and remaining in the retort, while the nitrous oxyde gas, or protoxyde of nitrogen, is driven over and collected in jars on the pneumatic cistern. The salt is a compound of the nitric acid and the alkali named ammonia. The carbonate of ammonia, in powder, is added to nitric acid, diluted with its own bulk of water, till the acid is neutralized; and this effect is manifested by dipping into the solution a slip of litmus paper, which is changed to green by an excess of alkali, and to red by an excess of acid. If the litmus paper be not at hand, paper, soaked in a decoction of red cabbage and dried, will answer; or, what is still more convenient, the blue paper wrapped round sugar loaves. The nitric acid unites with the ammonia, forming the nitrate of ammonia, and the carbonic acid, previously in combination with the alkali, escapes with effervescence.

I am in the practice of using the nitrate of ammonia in its crystalline state, and believe it to be a good plan. The retort should not be more than one third full, and is sufficiently heated by a spirit lamp, to insure entire decomposition. The salt speedily melts, and the gas is evolved freely. Some writers talk about the impurity of the first portions of the gas that come over; but if the salt be pure, I apprehend no danger on this score, especially as I do not employ it for the exhibition of its effects on the human system.

The theory of the production of the protoxyde of nitrogen from the nitrate of ammonia, is interesting, as all chemical theories are. We have here to anticipate, that the equivalent of nitric acid, in its absolute form, is 54. In the usual state, it is combined with two equivalents of water, 18, and hence the equivalent of the common acid is 72. The combining number of ammonia is 17, which, joined to 54, nitric acid, give 71 as the equivalent of dry nitrate of ammonia. This, by the action of heat, is changed into 27 parts of water, (3 equivalents,) and 44 of protoxyde of nitrogen or nitrous oxyde, (2 equivalents); and as 100 cubic inches of this gas weigh about 46½ grains, it follows, that 71 grains of the nitrate of ammonia will give 95 cubic inches of this gas, because, in their decomposition, 44 grains of nitrous oxyde are produced. This result is highly illustrative of the doctrine of definite proportions, and must be instructive to every one who desires to study the laws of chemical combination. For the more perfect understanding of the statement just made, the following diagram will be found useful.



In this diagram, we have, first, the equivalent of nitrate of ammonia, 71; next its proximate composition, viz. 54 nitric acid and 17 ammonia. Then comes the ultimate or elementary composition of nitric acid, viz. 5 equivalents of oxygen, making 40, and one equivalent of nitrogen, 14; after this, the elementary constitution of ammonia, viz. one equivalent of nitrogen, 14, and three equivalents of hydrogen, 3. Lastly, we have all these elementary principles entering into new combinations, and giving rise to three equivalents of water, 27, and two equivalents of the protoxyde of nitrogen, 44.

This gas supports combustion better than atmospheric air, because it contains half of its bulk of oxygen, in a condensed state; whereas, atmospheric air, according to the schedule

already exhibited, contains only one fifth of its volume of oxygen. If a candle be blown out, (the wick still remaining red,) and quickly dipt into a vessel full of the protoxyde of nitrogen, the flame will be revived. Phosphorus and sulphur burn in this gas, with great splendor, the latter with a rose-colored flame. These, and other combustibles, must be introduced in a state of active inflammation, because a higher temperature is required for the commencement, as well as continuance of combustion, in this gas, than in oxygen or atmospheric air. Exp.

We may form a detonating mixture with equal measures of this gas and hydrogen; placing them in a bottle or jar surrounded by a towel, and then applying a lighted taper. In this case, the hydrogen combines with the oxygen of the protoxyde, forming water, while the nitrogen remains.

This gas has been denominated *exhilarating*, from its peculiar action on the animal economy. When inhaled for a few minutes, it acts as a stimulus, in many cases, to the whole system. It often operates, in some respects, like ardent spirit, without leaving behind it the consequences of that poison. In some persons it excites wrathful feelings, in others the most pleasurable sensations, while in not a few, the effects are wholly deceptive. Of this I have been assured by some who have inhaled it, and who played the fool with tolerable success, but who have since declared, that they were simulating, all the while. Occasionally, it has done serious mischief, and laid the basis of incurable disease.\* I regard its administration to a class, as the very buffoonery of science, and think it should be made an indictable offence.

In all this, I am not to be understood as objecting to its use, as a remedial agent. For any thing I know to the contrary, it may prove a salutary article in the hands of judicious practitioners; indeed, it has been successfully employed in a variety of affections. One of our medical journals furnishes some satisfactory cases of pertussis, greatly relieved, and finally cured by exposure to an atmosphere of this gas; and I believe that it is entitled to further consideration.

It is made sufficiently pure for administration, as a remedial agent, by the process already detailed, and by allowing the gas to stand over water for one or two hours. This precaution, although it reduces the quantity of gas, by absorption, is favorable to the removal of nitrous vapors, if any should be present. If it support combustion well, there will be no reason to suspect its good quality. Its taste is somewhat sweet and pleasant, and it has an agreeable but faint odor. Before the patient attempts

\*The termination, in one instance, was mental derangement.

to inspire it, he should be directed to empty his lungs of common air, as far as may be practicable, by a deep respiration. Then the gas may be fairly breathed, five or six times in succession, taking care to close the nostrils at the same time. In order to have a perfect respiration of the gas, the tube connected with the gas bladder, should be of a capacious calibre.

A common bladder of large size, or an oiled silk bladder, will answer all the remedial purposes to which this gas may be applied.

Although the protoxyde of nitrogen may be inhaled for a time, without danger to the healthy subject, and even with benefit to the sick, yet experiments show, that an atmosphere of this gas would soon destroy animal life. Small animals placed in vessels filled with the nitrous oxyde gas, die in a very short period; the blood is found to be of a deep purple hue, and the irritability of the muscles is lost.

## SECTION II.—DEUTOXYDE OF NITROGEN.

This gas was noticed, at a very early period, by Dr Hales. It was examined with great care by Dr Priestley, about the year 1778, and by him called *nitrous air*. It is also known by the names, *nitrous gas* and *nitric oxyde*; but the more fashionable and correct appellation, is that which is placed at the head of this section.

The equivalent, by weight, of this gas, is 30, being composed of 14 (one equivalent) of nitrogen, and 16 (two equivalents) of oxygen; by volume, it is two measures, and is made up of one measure of nitrogen, and two half measures of oxygen; 100 cubic inches weigh 31.77 grains, and its specific gravity is, consequently, 1.0416. It is absorbed in small quantities, by water, and when it is mixed with common air, or oxygen gas, a dense, orange-red gas is formed, which is entirely absorbed by water.

The deutoxyde of nitrogen, (nitrous gas,) is usually prepared, by decomposing nitric acid by means of copper or mercury. For this purpose, copper clippings or filings are introduced into a tubulated retort, and nitric acid, diluted with twice its bulk of water is poured in. The metal being reduced to small particles, there is little or no force of cohesive attraction to be overcome, and hence we perceive a brisk action going on, as evinced by the effervescence. The liquid soon assumes a greenish blue color, and the copper is dissolved in the acid. During all this action, copious fumes of gas are given out, which may be collected in the usual manner. (See Fig. 1st, Plate 2d.)



Three equivalents of copper are required, in this process, for the complete decomposition of two equivalents of nitric acid. The copper takes the greater portion of the oxygen of the acid, and peroxyde of copper is the result; and as three equivalents of copper are used, so three equivalents of peroxyde of that metal are formed. As an excess of nitric acid is always employed, it combines with the peroxyde of copper, and hence we have the pernitrate of copper in solution. The portions of oxygen which do not combine with the metal, join the nitrogen of the decomposed nitric acid, and give rise to the deutoxyde of nitrogen. The following diagram offers a more satisfactory view of the action of the copper, on the nitric acid, than mere words could give.

1 equivalent nitric acid, =54	{	nitrogen	14	} 30 deutoxyde of nitrogen.
		oxygen	8	
		oxygen	8	
		oxygen	8	
		oxygen	8	
		oxygen	8	16+64 copper=80 peroxyde
		oxygen	8	of copper.
1 equivalent nitric acid, =54	{	oxygen	8	16+64 copper=80 peroxyde
		oxygen	8	of copper.
		oxygen	8	16+64 copper=80 peroxyde
		oxygen	8	of copper.
		oxygen	8	30 deutoxyde of nitrogen.
		nitrogen	14	
<hr/> 108 <hr/>			<hr/> 108 <hr/>	<hr/> 108 <hr/>

The excess of nitric acid, always employed in this process, is not brought into view, in the diagram, because there was no necessity for doing so, our object being to exhibit the decomposition of the two equivalents of acid, by the three equivalents of copper.

The solution of nitrate of copper, in the retort, may be readily crystallised by pouring it out into a saucer or plate, and as it rapidly absorbs moisture, it should be kept in close vessels.

The only metals that yield pure deutoxyde of nitrogen, by their action on nitric acid, are mercury and copper. The gas may be obtained from both metals, in the same manner, only that when mercury is used, very little heat is requisite, if the acid be not much diluted.

Deutoxyde of nitrogen is a transparent and colorless gas. We distinguish it, readily, from all other substances, by the dense, orange-red, suffocating fumes that are formed, when it comes in contact with oxygen gas, either pure. Exp.



or in combination with another gas. It is not a supporter of combustion, in the common sense of that term. A piece of burning sulphur or a common lighted candle is instantly extinguished, by immersion in this gas. But, if phosphorus or charcoal

**Exp.** be introduced into it, in a state of active combustion, they will burn with increased brilliancy.

These effects may be happily shown, by filling four jars with the gas. Into one, pass the lighted candle; into the second, the burning sulphur; place the lighted phosphorus in the third, and the ignited charcoal in the fourth, and the contrast will be very obvious.

The bodies which burn in this gas, effect its decomposition. Thus the phosphorus and charcoal unite with the oxygen, forming phosphoric and carbonic acids, while the whole of the nitrogen remains unaffected.

We are not able to form an explosive mixture by adding hydrogen to this gas, although it contains oxygen. To ascertain the truth of this statement, let equal portions (by measure) of these gases, be mixed in a long glass jar, and a lighted taper be applied to the mixture. We shall find that the hydrogen burns away with a silent flame, of a white color, tinged slightly with green. We infer from this result, that no chemical union ensues, when these gases are thus brought in contact.

We have already noticed the fact of a union between oxygen and the deutoxyde of nitrogen, and have stated the appearance attending that union. These gases evidently react on each other, but the nature of the resulting compound is modified by several circumstances. The relative quantities of these gases, in any mixture, have an agency in determining the distinctive features of the compound, so formed. But other circumstances also exert an influence; as, for instance, the size and shape of the vessel in which the mixture is made, the nature of the fluid over which they are mixed, the length of time they remain in contact, &c. &c.

**Exp.** Let an excess of the deutoxyde of nitrogen be mixed with oxygen, in a tube, containing a solution of caustic potash, over mercury. Every half measure of oxygen, (or one equivalent,) will combine with two measures of the deutoxyde of nitrogen, (or one equivalent,) and the caustic solution will absorb the whole of the gaseous mixture. The union of these gases constitutes an acid, (to be noticed presently,) called the hyponitrous acid, which is found combined with the alkaline solution, in the form of hyponitrite of potash. This acid must, necessarily, consist of one equivalent of nitrogen, or 14, and three equivalents of oxygen, or 24, making a total of 38. But

as it never has been obtained in a separate state, little is known about its peculiar properties.

Deutoxyde of nitrogen may also combine with oxygen, in the proportion of two measures of the former, (one equivalent,) and one whole measure of the latter, (two equivalents). If this combination be effected in a wide mouthed receiver, over water, a very considerable condensation results, and acid vapors are formed, which are rapidly absorbed by water. These are called nitrous acid vapors, and shall be further noticed.

Both of the acid compounds, just noticed, are formed, when equal volumes of common air and deutoxyde of nitrogen are mixed in a wide jar. One equivalent of the latter takes two equivalents of oxygen from the common air, which are condensed, and give rise to nitrous acid. On the other hand, two other equivalents of oxygen combine, each, with an equivalent of the deutoxyde, and two equivalents of hyponitrous acid are formed. Hence, as every measure of oxygen reacts on three times its bulk of deutoxyde of nitrogen, and a complete condensation takes place, on dividing the amount of this by four, we ascertain the quantity of oxygen which may be present in a given bulk of air. Guy Lussac has found, that the quantity of oxygen, in mixed gases, may be determined in the same manner, whether there be a larger or smaller proportion of oxygen present than exists in atmospheric air; taking care, always, to have a sufficient quantity of deutoxyde present, the condensation being always uniform, when this gas is added at once to the oxygen, in a wide vessel. When a narrow tube is employed for mixing the gases, the condensation is not uniform, and the resulting compounds are so slowly absorbed, that it is necessary to agitate them with the water, when a portion of deutoxyde is at the same time condensed.

We may make these experiments, by mixing air and oxygen with different proportions of the deutoxyde of nitrogen in glass jars, over the pneumatic trough, taking a small glass vessel to measure the different quantities. If an excess of deutoxyde be present, that can be shown, by the admixture of a little common air or oxygen with the gas supposed to contain it. And we can also know whether all the oxygen has been consumed, by introducing a small portion of the deutoxyde, by which ruddy vapours will be instantly formed. In this way, these gases are tests of the presence of each other; and by adding an excess of one to any gaseous mixture, containing the other, it may be entirely removed, because the resulting compounds are absorbed by water, with rapidity.

The solution of green muriate or sulphate of iron is blackened, when the deutoxyde of nitrogen is passed so freely through it, as to effect saturation. Sir Humphrey Davy employed this solution, to determine the quantity of oxygen in any gaseous mixture. It is not depended on, at present, however, because an evolution of deutoxyde commonly attends, or soon follows, the absorption of the oxygen gas.

The deutoxyde of nitrogen is quite irrespirable, exciting a severe spasm of the glottis, whenever an attempt is made to breathe it. There is great danger in an experiment of this sort, and no one should attempt it, for if the gas could reach the lungs, it would be changed into nitrous acid vapor, by the oxygen of the atmospheric air, and in that state would be highly irritating and corrosive.

### SECTION III.—HYPONITROUS ACID.

The equivalent of this acid has already been stated as 38, by weight, being a compound of 14 nitrogen and 24 oxygen. The method of obtaining it in combination with potash, has been noticed, and it is impossible to separate it from that union, so as to exhibit its distinct properties. If an acid be added to the compound, it combines with the potash, and the hyponitrous acid, although detached, is resolved into nitrous acid, and deutoxyde of nitrogen.

As this is the first acid body we have been called to notice, it is proper, before proceeding further, to make a few remarks on the general properties of acids.

Acids, for the most part, are sour to the taste, and soluble, in a greater or less degree, in water. They redden most of the vegetable blues, and restore the colors that have been changed by alkalies. They combine with all the metallic oxydes and form salts. Some of them are corrosive. They exist in the solid, fluid, and gaseous states. All these marks, taken together, may serve for a definition of an acid.

As to the *acidifying principle*, I hold the same sentiment now, that I advocated twenty years ago; and that is, that such a principle is the creature of fancy, especially when identified with any single agent. I have maintained the position in the *New York Medical Repository*, and the *Memoirs of the Columbian Chemical Society*, so far back as 1812–13, that acidification, like combustion, is the result of relative actions, and can never, philosophically, be attributed to any one substance. In short, all the elements, essential to any acid, are alike entitled to the

appellation of acidifying principles. On the very same ground, I reject the terms, principle of inflammability, and principle of alkalinity.

## SECTION IV.—OF NITROUS ACID.

The equivalent of this acid, by weight, is 46; by volume, one measure. It is composed of 14, or one equivalent of nitrogen, and 32, or 4 equivalents of oxygen. The specific gravity of the liquid acid is 1.452, while that of the vapors is 3.19. One hundred cubic inches weigh 97.428 grains.

The following is the usual method for preparing this acid, by decomposition. The substance called nitrate of lead, (composed of nitric acid and lead,) is exposed to heat in a green glass retort, placed in a sand bath, till it is completely decomposed, and the disengaged vapors are condensed in a receiver, kept as cold as may be practicable. In this process, the nitric acid, previously in combination with the lead, is resolved into nitrous acid and oxygen; the former being condensed into an orange-red colored and fuming liquid, while oxyde of lead remains in the retort. Exp.

We may form the same acid, in the gaseous state, by mixing two measures (one equivalent) of deutoxyde of nitrogen with one whole measure (two equivalents) of oxygen. This mixture should be made in a flask previously exhausted by the air-pump, and a condensation will be found to take place in the mixed gases, equal to two-thirds of their whole volume. In this way, it is determined that the equivalent of nitrous acid, by volume, is one measure. The dry gases, thus mixed, form nothing but nitrous acid vapors, and they must not be combined over either water or mercury, as the former would render the production of nitric acid certain, while the latter would decompose the nitrous acid.

If we expose liquid nitrous acid to the air, evaporation goes on rapidly, and the vapor is highly corrosive and suffocating. This vapor will support the combustion of a taper or phosphorus, but it will extinguish sulphur and other inflammables, although introduced in a state of active inflammation. Exp. A large quantity of water will decompose it, completely, forming nitric acid and deutoxyde of nitrogen; the former combines with the water, and makes a colorless solution, while the latter escapes with effervescence. If a smaller quantity of water be employed, the acid vapor is changed in color, from a deep reddish-brown to a greenish-blue; and at last, it becomes



quite colorless. This acid and its solutions redden vegetable blues, and they are highly acrid and corrosive.

It is only for experimental illustrations, that nitrous acid is ever prepared in a pure state. It is a powerful oxydising agent, however, and might be employed for this end, as it readily imparts its oxygen to many bodies. When transmitted through red hot porcelain tubes, it is decomposed; oxygen and nitrogen gases are the products.

#### SECTION V.—NITRIC ACID.

This acid exhibits the highest state of oxygenation of which nitrogen is susceptible. Its equivalent is 54, being composed of one equivalent of nitrogen, 14, and five of oxygen, 40. This denotes pure nitric acid. The common liquid acid contains two equivalents of water, 18, which added to 54, the equivalent of the pure acid, give 72 as the number for the common acid. It is transparent and colorless, but soon changes to straw color, by the action of light. It gives out suffocating fumes, on exposure to the air, and readily absorbs water. It is intensely acid and caustic. Its specific gravity is 1.500; boils at  $210^{\circ}$  and congeals at  $62^{\circ}$  below zero.

This acid was known so far back as the time of Raymond Lully. Basil Valentine describes the process for making it, and calls the product *water of nitre*. It has since been called *spirit of nitre* and *aqua fortis*.

Absolute nitric acid has not been obtained in an insulated state, and the term nitric acid is usually applied to the compound of water and acid, just referred to. This article is ordinarily prepared, by decomposing the common saltpetre by means of sulphuric acid. A tubulated retort is preferable to a plain one, for this purpose, because we can more easily introduce the materials, than if a plain retort be selected. If the latter be employed, the acid must be passed in by means of a funnel, long enough to reach into the body of the retort, and thus avoid the lodgment of any portion of acid along the beak. The bruised saltpetre, with an equal weight of sulphuric acid, being placed in the retort, should be exposed to heat in a sand bath, or over a spirit lamp or chauffer, and the retort should not be more than one third full, because there is danger of the materials being driven over.

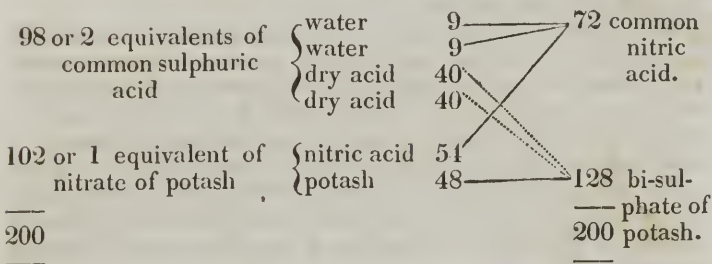
The acid, as it is separated from the saltpetre by the decomposing energy of the sulphuric acid, passes over into a receiver closely fitted to the retort, and kept cool by a stream of cold



water constantly falling on it. This condenses the acid, and we may continue the distillation, till red vapors have been, for some time, visible in the retort.

Light green glass retorts should be preferred to those of flint glass, for this operation, because they are less liable to be broken, and do not so easily melt.

In explaining this process, it should be recollected, that salt-petre is composed of nitric acid and potash, and that the sulphuric acid merely detaches the potash from this connexion, and sets the nitric acid free. The materials are mixed so nearly in the proportion of two equivalents of sulphuric acid to one of nitrate of potash, that we may assume this to be the case in explaining the reaction which ensues. Every two equivalents of the common sulphuric acid consist of two equivalents of water, 18, and two of real sulphuric acid, 80, making in all 98; while the nitrate of potash, is composed of one equivalent of potash, 48, and one of nitric acid, 54, making 102. The dry, or real sulphuric acid combines with the potash, forming sulphate of potash; but as two equivalents of acid unite with one equivalent of potash, we term the product, *bi-sulphate* of potash. The dry nitric acid liberated, combines with the water of the sulphuric acid, and forms the liquid condensed in the receiver. The annexed diagram will give a more satisfactory exhibition of the process.



The proportions we have given for preparing this acid, are such as are used by the London college, and they are recommended by Gray, in his 'Operative Chemist.' The Edinburgh and Dublin colleges take two parts of sulphuric acid to three of nitre. One equivalent of sulphuric acid is sufficient to decompose an equivalent of nitrate of potash, but as it contains only one equivalent of water, and as the nitric acid of the nitre requires two equivalents, we could not procure this acid, unless more sulphuric acid were employed. As fast as we should separate the nitric acid from the potash, it would be decomposed into nitrous acid and oxygen.

As we have already observed, the nitric acid of the nitre, comes over with the water of the sulphuric acid; but towards the end of the process, the nitric acid is decomposed, and red, dense fumes of nitrous acid are evolved, which, for the most part, are condensed by the nitric acid in the receiver, if it is kept sufficiently cool.

Nitric acid may also be formed, by passing a succession of electric sparks through a mixture of oxygen and nitrogen gases. It was in this mode, that the acid was first obtained by Cavendish, in 1785.

An impure nitric acid was made, formerly, from nitre and copperas; the former being dried and the latter calcined to redness, before mixture. The whole was exposed to heat, and the acid vapors collected in a suitable receiver.

The best method, however, is that in which nitre and sulphuric acid are employed. We may show this process, on a small scale, by taking three ounces, by measure, of sulphuric acid, and eight ounces, by weight, of coarsely powdered nitrate of potash, and put them in a retort of such a size as that it shall be rather less than half filled by the materials. Expose the retort to a lamp heat, and continue the distillation till a quantity of acid shall have been obtained, equal in measure to the sulphuric acid employed. A further continuance of heat, would enlarge the product a little. During the operation, cold should be applied to the receiver.

By whatever process nitric acid is prepared, it has usually a slight tinge of yellow, owing to the presence of some nitrous gas, formed by the decomposition of a portion of nitric acid, during its preparation. To free it entirely of color, it must be exposed to a gentle heat, as long as any nitrous acid vapors, or deutoxyde of nitrogen are expelled. If nitric acid be exposed to the light, there will be a partial decomposition, with evolution of oxygen gas, and formation of nitrous acid. Thus, a variety of tints will be imparted, from a light straw color to a deep orange.

Unless we desire to have a sample of colorless nitric acid, it is hardly necessary to expose the mixed acid to heat, with a view to drive off the gaseous matters which produce discoloration; because it is equally useful, for almost all purposes, in the apparently impure form. It is even preferred for many uses, because it is stronger and more active than the pale acid. Dr Hope has prepared it with so high a specific gravity as 1.54, while the specific gravity of the colorless acid does not exceed 1.500.

It is proper to say, further, in regard to the manufacture of

nitric acid on the large scale, that a stoneware apparatus is usually employed, analogous to the ordinary glass retort and receiver, and that a less quantity of sulphuric acid is used than that which we have designated.

A pretty good evidence of the strength of nitric acid, for all pharmaceutical uses, is furnished by its power to neutralise a given quantity of crystallised carbonate of soda. One hundred grains of nitric acid, as pure as it can be made in the usual way, will neutralise 212 grains of this salt.

Nitric acid attracts water from the air, and unites with it in all proportions, and the combination is attended with a considerable degree of condensation or evolution of heat. According to Dr Ure, the greatest amount of condensation occurs when 58 parts, by weight, of acid, (specific gravity 1.500,) are mixed with 42 of water. The bulk is reduced one-twelfth, and the temperature rises about 80°.

What is known in the arts and in commerce, by the term *aqua fortis*, is a diluted nitric acid, prepared by distillation from nitre, with diluted sulphuric acid. It is estimated to contain only one-fourth part as much acid, as the strong nitric acid. *Double aqua fortis* contains twice as much strength as the aqua fortis, No. 1, being half as powerful as the best nitric acid. *Treble aqua fortis* is another variety, which is known to some artists.

The diluted acid of the Edinburgh and Dublin colleges, consists of equal parts of the strong acid and water, by weight; but one is led to conclude, that this dilution was designed only for chemical purposes. A much more suitable preparation for medicinal use, is that of the London college. It is made by adding one measured ounce of acid to nine ounces of water.

The quantity of water mixed with this acid, determines its boiling point. If its specific gravity is 1.42, it boils at 248°, and distils unaltered. Its freezing point is influenced in a similar manner. At the specific gravity 1.420, a degree of cold equal to 41° below zero, is requisite to congeal it. When this acid is poured upon snow, the latter speedily melts, and an intense degree of cold is produced. The sudden solution of the snow demands all the latent caloric of the acid, which, being again latent in the dissolved snow, a reduction of temperature necessarily follows.

If nitric acid be exposed to a red heat, in a porcelain tube, it is resolved into oxygen and nitrogen gases, which may be collected in a pneumatic trough. To perform this experiment, take a tube about three-fourths of an inch in diameter, stuffed with fragments of earthenware, (in order to enlarge the internal surface) and pass it through a furnace or chauffer. To one

end of the tube, attach a small tubulated glass retort, and to the other a bent glass tube, and let the extremity of the bent tube dip under the shelf of the pneumatic cistern. Then lute the apparatus together, with a mixture of wheat flour and water, white of egg and fine lime. The porcelain tube being made red hot, let a small quantity of nitric acid be poured into the retort, and boiled by means of a spirit lamp, having the stopper of the tubulure removed until ebullition actually takes place. If the stopper were to remain in its place from the first, there would be risk of a vacuum in the retort, and consequently of a rush of water into the heated tube, by which the latter would be broken. As soon as boiling occurs, the tubulure is closed, the acid vapors are driven over, and coming in contact with the red hot tube, are decomposed into oxygen and nitrogen gases, which may be collected in suitable jars. Care must be taken to heat the tube gradually, as there is some danger of breaking it.

Nitric acid emits very acrid fumes, when exposed to the air. It exhibits, in an eminent degree, the characteristic features of an acid, corroding animal and vegetable substances, reddening vegetable blues and forming salts, called *nitrates*, by union with the different salifiable bases. Its taste is intensely sour, even when largely diluted with water, and it stains the skin of a yellow color, which continues till the cuticle disappears. It is decomposed by most metals and combustible bodies, yielding oxygen to them, with a readiness that is peculiar to it. In these cases, however, it does not impart the whole of its oxygen; enough remains in combination with nitrogen, to give rise to an evolution of deutoxyde of nitrogen and nitrous acid.

Notwithstanding the facility with which this acid gives out its oxygen to certain bodies, it exerts no action on some substances, when in a concentrated form. Professor Woodhouse, was the first to remark, that pure nitric acid has no action on metallic copper.\* If water be added, the work of decomposition soon commences, and the metal takes oxygen, both from the acid and the water. Under these circumstances, there is necessarily an evolution of hydrogen, which, uniting with a portion of nitrogen, given out in the decomposition of the nitric acid, forms ammonia; and this, combining with some undecomposed nitric acid, gives rise to a small quantity of nitrate of ammonia. This statement is useful, by explaining the source of the white fumes which are sometimes seen, closely blended with the nitrous acid vapors that are formed, when nitric acid is decomposed by a metal that has a strong affinity for oxygen.

\* See Woodhouse's edition of Chaptal's Chemistry, vol. II.



We have said, that the colored appearance of nitric acid is dependent on the presence of the deutoxyde of nitrogen; and this can be shown to be the case. If a stream of the deutoxyde be freely passed through colorless nitric acid, a large portion of the gas will be absorbed, the acid will acquire a light straw hue, deepening into reddish-brown, and passing through various shades of olive and green, till at last it becomes partially blue. If we expose the acid, thus discolored, to the action of heat, there will be a disengagement of nitrous acid vapor and protoxyde of nitrogen, and the liquid will exhibit its original appearance. The apparatus best adapted to this purpose, (because the most economical and easily fitted up,) is a Woulfe's bottle, an oil flask, and two bent glass tubes. The middle aperture of the bottle is to be closed. The tube proceeding from the flask, is to enter one of the openings in the bottle, and from the other opening, another bent tube is to be conducted to the penumatic tub, to carry off any excess of gas. The materials for making the deutoxyde of nitrogen are to be put into the flask, and a quantity of pale nitric acid in the Woulfe's bottle. Exp.

In this process, there is evident decomposition. The deutoxyde takes oxygen from the nitric acid, converting itself as well as the acid into nitrous acid, and the intensity of the color produced in the acid contained in the bottle, depends on the quantity of deutoxyde absorbed. In the estimation of some chemists, a small portion of hyponitrous acid is formed, but this is not fully ascertained.

What is sometimes called *strong fuming* nitric acid, is usually a compound of nitric and nitrous acids. It is prepared by distilling nitre with two thirds of its weight of sulphuric acid. If small quantities of water be added to it, the orange-red color is gradually changed to olive and green. If still more water be added, the whole is made colorless.

From carelessness in conducting the distillation of nitric acid, it is sometimes contaminated with sulphuric acid; and as nitre often contains muriatic salts, there is sometimes an adulteration with muriatic acid. We detect these acids by means of the tests called muriate of barytes, and nitrate of silver. The nitric acid, to be examined, should be first diluted with about four parts of water. To one portion of the diluted acid, add a few drops of muriate of barytes, and a white precipitate will be instantly formed and soon collected at the bottom, which consists of the barytes and the sulphuric acid, which happened to be present. These are now combined, in the form of sulphate of barytes. To another portion of diluted acid, we add some nitrate of silver to detect muriatic acid. The muriatic acid having an affinity for



silver, greater than nitric acid has, combines with it, and a curdy precipitate, of a greyish white, is formed, which is a muriate of silver.

To purify nitric acid from the sulphuric, it must be redistilled, with an additional quantity of nitre, (nitrate of potash). The sulphuric acid decomposes a portion of the salt, and is, of course neutralised by the potash. To separate muriatic acid, the nitrate of silver must be added as long as any precipitate is formed, and then pure nitric acid may be obtained by distillation.

We distinguish nitric acid, when pure, by the facility with which it is decomposed by inflammable substances, among which we include several of the metals, as zinc, tin, mercury, and copper; also, by the large quantity of deep, ruddy-colored fumes which are evolved, and by the peculiar odor of those fumes. But it is not so easy to determine its presence when in minute quantities, as there are no substances with which it forms characteristic and insoluble precipitates. A test has been proposed by Dr Liebig, but I am not disposed to place much reliance on it. He directs the liquid under examination, to be mixed with a solution of indigo in sulphuric acid, till it acquires a perceptibly blue color. Then add a few drops of sulphuric acid to the solution, and boil the whole for a short time. If the liquid contain any nitric acid, it is said that it will immediately lose its color, or if the proportion of acid be very minute, the fluid will be rendered yellow. It is affirmed, that one part of nitric acid may thus be detected in four hundred parts of water.

I know of no better method by which to determine the presence of nitric acid in any suspected solution, than the following: Add a small quantity of sulphuric acid, in order to separate the base, which may, probably, be united to a portion of nitric acid, and then drop in a little muriatic acid, together with a small piece of gold leaf. If any nitric acid be present, it will instantly form a compound with the muriatic acid, which will dissolve the gold. This test, however, is not infallible, since chloric acid (which, by the way, is rarely present) will exert the same action on the metallic leaf.

As a general rule, we may affirm, that if on adding gold leaf to a solution supposed to contain nitric acid, the leaf is not acted upon, but is quickly dissolved by pouring in some muriatic acid, we may infer that nitric acid was really present.

With one or two exceptions, the salts of nitric acid are quite soluble in water, easily decomposed by heat, and deflagrate with inflammable substances. Sulphuric acid is capable of detaching nitric acid from all its combinations, if sufficient heat be employed.

To show how readily combustibles decompose nitric acid, put a few drachms into a long glass tube; then throw in some small pieces of phosphorus, and in a few minutes combustion will ensue. If the effect does not take place instantly, <sup>Exp.</sup> dip the lower end of the tube in hot water. The phosphorus takes oxygen from the acid, and dense vapors of phosphoric acid are formed, nitrous fumes being evolved at the same time.

Owing to a property of nitric acid to which we have alluded, viz. the facility with which it yields oxygen to other bodies, it is much employed in chemical processes, as in the formation of several metallic nitrates. It is also employed in the process of etching on copper, making *aqua regia*, (the solvent of gold,) and in preparing sulphuric acid, nitric ether, &c.

Nitric acid vapor has been employed for the purposes of fumigation, in order to destroy the matter of contagion. For this purpose, an ounce of powdered saltpetre is directed to be placed in a teacup, and the same quantity of sulphuric acid poured on it. The cup is then to be placed in a bason nearly full of hot water, to excite decomposition. In this way, copious fumes will be disengaged for several hours. But it is believed that a much better fumigator can be employed, viz. chlorine, of which more will be said, in this relation, hereafter.

In the diluted state, nitric acid is employed as a tonic, and by some physicians is highly extolled in the treatment of the consecutive effects of opium. Externally, it has obtained some celebrity in medical practice, in the form of the nitro-muriatic bath, applied, generally, to the feet and legs, by means of a sponge. The compound acid is made by mixing two parts of muriatic with one of nitric acid, and then adding so much water as to make the mixture taste as sour as very strong vinegar, and to produce a slight prickling sensation on the skin.

Strong nitric acid is also used as an escharotic, and has been successfully applied to sloughing, phagedenic ulcers.

The doctrine of *incompatibles* has some application to nitric acid, although by reason of the violent action occasioned on mixing it with many compounds, the most ignorant might be sufficiently warned. We mean, generally, by an incompatible medicine, that, which on coming in contact, or mixing with some other medicine, given for a special purpose, either greatly alters or entirely destroys its remedial character, and sometimes gives rise to a deleterious compound. If we give nitric acid to obtain its tonic effects, it would be incompatible to administer with, or directly after it, either magnesia or potash, or the blue

(mercurial) pill. The two first would greatly interfere with the object in view; the latter would jeopardise the patient's life.

A compound, similar in some respects to oxygenated water, (or the deutoxyde of hydrogen) has been formed, by dissolving the deutoxyde of barium in diluted nitric acid. The barytes is then to be separated from the nitric acid by adding sulphuric acid. An insoluble sulphate of barytes will be thrown down, from which the clear liquor is to be decanted and then made stronger, by evaporation in the exhausted receiver of an airpump. The rationale agrees with that we gave, when treating of oxygenated water. The deutoxyde of barium contains two equivalents of oxygen, one of which attaches itself, somehow or other, to the nitric acid, leaving the protoxyde of barium, (or barytes,) to unite with the sulphuric acid. The general relations of oxygenated nitric acid and oxygenated water, to metals and metallic oxydes, are alike.

It is proper to remark here, that many experiments fail in the hands of the young chemist, although properly conducted, as to the manipulations, owing to defect in the materials. The same thing occurs, occasionally, in the operations of older men, but they generally understand the difficulty. Now, there is no article which more frequently leads to disappointment, than nitric acid; and hence the importance of preparing a sufficient quantity of this article in the laboratory, to answer for all nice and delicate experiments.

It has been well said of fire and water, 'that they are eminently good servants, but most terrible masters,' and the same remark applies, with great propriety, to nitric acid. It is highly useful in the arts, in chemistry and in practical medicine; but an overdose, either by mistake or design, is almost invariably productive of fatal results. It is important, therefore, to know, by what means the poisonous tendency of this article may be most happily counteracted; and chemistry, alone, can direct us in this matter. Some agent must always be employed, if practicable, that will neutralise the poisonous matter, and form, by combination, a comparatively harmless substance; and herein consists the grand secret of preventing the baneful effects of poisons, so far as chemistry is concerned. On this principle, Orfila and others have very properly directed the immediate use of calcined magnesia, not only for nitric acid, but for all the mineral acids. One ounce of magnesia must be diffused in a quart of water, and a half pint of the mixture given every two or three minutes. If this article be not at hand, let flaxseed tea or water be administered in large draughts, so as to dilute the acid and sheathe the stomach. A half ounce

of soap, dissolved in a quart of water, will also be a proper mixture, and should be passed into the stomach without delay, if the magnesia cannot be had instantly. The soap acts by virtue of the alkali it contains, which combines with the nitric acid, and forms a comparatively mild salt. As a substitute for all the above, finely powdered chalk may be suspended in water, and given liberally. In such cases, not a moment is to be lost, for delay will insure fatal results. Injections, of the various articles given by the mouth, may be freely employed, as they are calculated to have similar effects. After these means shall have succeeded, so far as it possible for them to be useful, the patient is without the province of chemical science, and is a proper subject for that medical treatment which is demanded by the consecutive effects of all poisonous matters.

#### SECTION VI.—ATMOSPHERIC AIR.

This substance has been placed here, because, as will be seen, its composition does not exactly harmonize with the law of equivalents. By weight, it is called 36; consisting of two equivalents of nitrogen, 28, and one equivalent of oxygen, 8. It presents the first instance in the combinations of nitrogen, in which that element has been known to combine in more than one proportion. And if we were to preserve the uniformity in regard to it, and say that atmospheric air was composed of 14 nitrogen and 4 oxygen, then we would offer an anomaly in relation to the latter element.

The specific gravity of atmospheric air, being the standard of comparison, in estimating the specific gravity of other gases, it is reckoned 1 or 1000. One hundred cubic inches weigh 30.5 grains; and it is 828 times lighter than water.

Scheele and Lavoisier are the reputed discoverers of the composition of atmospheric air. Both experimented at the same time, or very nearly, but without concert. They determined, however, nothing more than the nature of the component parts, leaving the relative proportions of the oxygen and nitrogen to be settled by other chemists. We have had occasion, when treating of nitrogen, to notice an experiment, which seems to have been the original process of Scheele, in this investigation. He poured two ounces of a solution of the sulphuretted hydrosulphuret of potash or lime into a bottle, capable of holding six times that quantity. The bottle was accurately closed with a glass stopper, and opened under water after the lapse of about two weeks. This length of time is unnecessary, if frequent agitation be employed. Let the bottle be well shaken for one hour, and it will be evident, by the

Exp.



sudden rush of water, to supply the partial vacuum, as soon as the stopper is removed, that the oxygen of the contained air has been absorbed. The gas, remaining in the bottle after this operation, is nitrogen.

There is an uniformity in the composition of atmospheric air, in all regions, that is indicative, in a high degree, of the wisdom and goodness of the great Creator. It is this, that has fitted all lands for the abode of man and inferior animals, and thus constitutes us citizens of the world. Some inaccurate experiments led for a while to the belief, that the proportion of oxygen in the air of the atmosphere, was not always and every where the same. This subject, however, has been carefully investigated, and it is now settled, that in all parts of the globe, (excepting here and there a peculiar location, marked by gaseous emanations,) at all elevations and in all seasons, the proportions of oxygen and nitrogen are the same. The contrary opinion led to the use of the instrument called the *eudiometer*,\* of which there are now several varieties in use. On Plate 2d, Fig. 7th, may be seen the more usual kind of this instrument, although it is there explained with reference to the detonation of gases. The eudiometer was formerly employed, simply to ascertain the relative oxygenation of atmospheric air, in different places. At present, we apply it to estimate the proportion of oxygen in all the mixed gases.

We have already stated the method by which hydrogen and deutoxyde of nitrogen enable us to determine the quantity of oxygen gas, in any æriform mixture. It has been shown lately, by Dr Thomson, that on detonating hydrogen with atmospheric air, 100 parts, by measure, of the latter, must be mixed with 42 measures of the former, in order that all the oxygen may be consumed. If less hydrogen be employed, a portion of oxygen will remain, and if more be used, some of it will unite with the nitrogen of the atmospheric air, and form ammonia, by which a greater condensation will be effected, than would otherwise occur.

Phosphorus is also employed for eudiometrical purposes. It has a natural tendency to combine with oxygen, and when no opposing circumstances intervene, the combination cannot fail to take place. A stick of it is to be introduced into a jar or tube containing the air about to be examined. The vessel is  
 Exp. first filled with water, then inverted on the shelf of the pneumatic cistern, and the air or gas is passed up. The water being thus displaced from the tube, a stick of phosphorus is easily introduced. The white vapors of phosphorous acid are

\* This word means, literally, a *measurer of air*.



soon perceived, resulting from the union of the phosphorus with the oxygen present in the tube. These vapors are speedily condensed and absorbed by the water, which rises in the tube to fill the partial vacuum. If the air thus operated upon be common air, there will be nothing left in the tube but nitrogen, excepting a very minute portion of the vapor of phosphorus. To insure the success of this experiment, the tube should remain undisturbed, for at least twelve hours; and, in all cases, the instrument should be divided into one hundred equal parts.

A small quantity of watery vapor and carbonic acid are always present in atmospheric air. The former may be detected by potash, which absorbs the moisture and becomes liquid; the latter, by lime water, the lime of which unites with the carbonic acid, and forms an insoluble carbonate of lime. To show these results satisfactorily, put some dry potash Exp. on a clean plate, and a quantity of fresh, transparent lime water, into a broad, shallow dish, and let these be exposed to the open air.

The active chemical agency of atmospheric air, is dependent, chiefly, if not entirely, on the oxygen which it contains. This is evinced in combustion and respiration, as well as in those chemical changes that are constantly occurring at the earth's surface, where atmospheric air acts an important part.

Astronomical observations teach, that the atmosphere is about forty-five miles in height, and that its specific gravity lessens, in proportion to the distance from the surface of the earth; because the smaller the superincumbent column, the less the pressure, and of course the density must be less. Elasticity and compressibility belong to atmospheric air, in common with other gaseous fluids. Mr Perkins alleges, that it may be liquefied by a pressure equal to 2000 atmospheres.\*

The compound blowpipe has been noticed in this volume, at some length, and it is proper here to speak of the common mouth-blowpipe. This may be made of glass or metal. It is, usually, from six to ten inches long, in the form of a hollow tube, the large end of which is from a fourth to a half inch in width, and there is a gradual tapering to the small end, the aperture of which is very minute. The mouth embraces the large end, and the operator forces the air from his lungs through the tube, directly on the flame of a candle or lamp, and thus maintains an elevated temperature, which is applicable to many useful purposes. In this instrument, it is obvious that atmospheric air plays an important part.

\* Or 30,000 pounds to the square inch.

In addition to the reason already assigned, for placing atmospheric air at the end of the list of substances composed of nitrogen and oxygen, I may add, the unsettled question respecting its actual character. Is it a simple mixture, or a chemical compound? Now, although I am unwilling to pronounce a positive judgment in this matter, I feel at liberty to say why, in my opinion, this substance is formed on chemical principles, rather than by mere mechanical mixture.

And here it may be affirmed, that in the operations of chemical attraction between any two bodies, an entire change of properties follows. What, then, are the properties of oxygen and nitrogen in a separate state; and, what are the peculiarities of atmospheric air? If none of the essential and characteristic qualities of nitrogen are found in atmospheric air, then it follows, that a chemical action has ensued, between it and oxygen, in the formation of common air. The leading feature of nitrogen is, unfitness for respiration and combustion; animals die when subjected to its influence, and lighted tapers are instantly extinguished. But atmospheric air is the grand element of life; and equally indispensable in the maintenance of ordinary combustion.

There are also dissimilarities between oxygen gas and common air. The former is too stimulating to be constantly inhaled, while the latter is precisely adapted to the condition of the lungs in the healthy state. They differ moreover in specific gravity, and in their powers to maintain combustion. It is plain to the most cursory observer, that the distinctive features of nitrogen and oxygen do not exist in atmospheric air.

But it is said, that nitrogen and oxygen will be found detached from each other, in a tube that has been filled with common air, and closed perfectly tight, after some months have elapsed; and so, also, will mercury be separated from the oxygen with which it was combined in the form of oxyde, after the lapse of a given period; and yet no one, for this reason, denies that oxydes are chemical compounds.

These are the views that I have entertained, on this subject, for twenty years; for a fuller account of them, the reader is referred to the *Memoirs of the Columbian Chemical Society*, published at Philadelphia, in the year 1813.

The next article to be noticed, is a compound of nitrogen and hydrogen, denominated *ammonia*, a word derived from *Jupiter Ammon*, in Africa, near whose temple, the salt called *sal ammoniac* was originally found. As we usually meet with this substance in the form of gas, it is very frequently denominated *ammoniacal gas*. It is regarded as a compound of 14

(one equivalent) nitrogen, and 3 (three equivalents) hydrogen, making its combining number, 17. But as I incline to the views of Berzelius and others, in relation to the metallic base of ammonia, I shall defer the full consideration of this subject, to the chapter which treats of the alkaline metals.

---

## CHAPTER VI.

### Of Sulphur.

PROCEEDING on the principles of combination, based on the law of definite proportions, the equivalent of sulphur, by weight, is ascertained to be 16, and by volume, one measure. Experiment has shown, in relation to all the well defined compounds of sulphur, with oxygen and hydrogen, that this estimate is correct. The usual specific gravity of this simple body, is 1.99; but if we melt it by means of a heat of  $450^{\circ}$ , and then cast it into water, its density is increased to 2.325.

Sulphur is a solid, inflammable substance, of a light yellow color, very brittle, with little or no taste, and emitting a peculiar odor, when rubbed. It is an abundant production of the mineral kingdom, and is found in large quantities in volcanic countries, as Sicily, Italy, and Iceland. It is frequently obtained in pyramidal crystals, and when melted and cooled slowly, it always presents a crystalline structure. Sulphur is also found in combination with a number of the metals, as iron, lead, copper, and antimony; and by exposing the common yellow iron pyrites, (bi-sulphuret of iron,) to a red heat in close vessels, it may be procured in considerable quantity.

When exposed to heat, it melts at the temperature of  $216^{\circ}$  or  $220^{\circ}$ , and becomes very thin and fluid at  $250^{\circ}$ . The roll sulphur of commerce, is prepared by pouring the melted sulphur, in this state, into cylindrical moulds; it gives a crackling noise when held in a warm hand, and often falls to pieces, having little cohesion, and being unequally expanded by the heat, as it is a bad conductor of caloric. When exposed to a stronger heat, instead of becoming more thin and fluid, as is generally the case with other liquids, it soon begins to turn thick and viscid, and at the temperature of  $450^{\circ}$ , the vessel containing it may be inverted without any of the sulphur falling out. By a farther increase of temperature it again becomes more fluid, and if it is then poured into water, it becomes a soft and tenacious mass, which may be used for taking impressions of seals or

medals, becoming hard after some time has elapsed. All these circumstances may be easily seen, by heating sulphur in a dry Florence flask, over a chauffer, or Argand's lamp, and supporting it on a retort stand.

To obtain large crystals of sulphur, the best method is to melt one or two pounds in an earthen crucible, and to invert it after it has been removed from the fire and a crust formed on the surface, a small hole having been made, to allow the sulphur, that is still liquid, to flow out. When the crucible is quite cold, it may be broken, and crystals of sulphur will be found lining the cavity previously occupied by the sulphur which had been poured out.

Sulphur is converted into vapor at  $600^{\circ}$ , and may be condensed, unchanged, by conducting the sublimation in close vessels. This process is generally resorted to for the purpose of purifying common sulphur, and when the heat is not too great, the vapor of the sulphur condenses in very minute crystalline grains, forming what is commonly called *flowers of sulphur*. To show the sublimation of sulphur, it may be exposed to heat in a retort with a short neck, connected with a large receiver. This article, also called *sublimed sulphur*, is occasionally contaminated with sulphurous acid gas. To get rid of this, it is necessary to wash it freely with water; especially when it is used for medicinal purposes.

Sulphur takes fire at about  $300^{\circ}$ , when heated in the open air, burning with a blue flame, and producing very pungent, suffocating fumes. In oxygen gas, it burns much more vividly and with a larger flame; sulphurous acid gas is the product of the action, in both cases, unless when the gases are moist, a small portion of sulphuric acid then being formed; the oxygen neither increases nor diminishes in volume, but becomes twice as heavy as before, from the sulphur which it takes up.

Sulphur is insoluble in water, but combines with it when precipitated from any solution containing it, forming a white powder, usually termed the *milk of sulphur*. The method of preparing it will be described in the last section of this chapter; it is the *precipitated sulphur* of the pharmacopœia. Sulphur combines with alcohol, when they are presented to each other in the gaseous state; it is sparingly soluble in ether, but is dissolved readily when boiled in oil of turpentine.

Sulphur combines with the metals, forming a great number of very important compounds, and during the combination, heat and light are frequently disengaged; a certain elevation of temperature is almost always necessary, however, to commence the action, and the appearance of the product varies considerably, according to the temperature at which it takes place.



If iron, in the state of ignition, or, still better, at a white heat, be rubbed with a roll of sulphur, the iron melts and falls in drops of liquid sulphuret. In this way, a bar of iron may be divided. In accordance with this fact, is an experiment of Professor Hare, as follows: Heat the but-end of a gun-barrel to redness, or rather to a white heat; then throw in at the muzzle a piece of sulphur, and blow the vapor out of the touch-hole. If a bunch of iron wire be held opposite the touch-hole, at this juncture, it will be burnt, as if in oxygen gas, and fused globules of protosulphuret of iron will fall to the ground.

The annexed table represents the constitution of the compounds of sulphur with oxygen and hydrogen, and which will be considered in the following sections.

	Sulphur.	Oxygen.		Sulphur.	Oxygen.
Hyposulphurous acid	32 + 8 = 40	or	<input type="checkbox"/> <input type="checkbox"/>	+ <input type="checkbox"/>	= ?
Sulphurous acid	16 + 16 = 32	or	<input type="checkbox"/>	+ <input type="checkbox"/>	= <input type="checkbox"/>
Sulphuric acid	16 + 24 = 40	or	<input type="checkbox"/>	+ <input type="checkbox"/> <input type="checkbox"/>	= ?
Sulph. acid. Water.					
Liquid sulphuric acid	40 + 9 = 49				
Sulphuric acid. Sulphurous acid.					
Hyposulphuric acid	40 + 32 = 72				
Sulphur. Hydrogen.					
Sulphuretted hydrogen	16 + 1 = 17	or	<input type="checkbox"/>	+ <input type="checkbox"/>	= <input type="checkbox"/>
Supersulphuretted hyd.	32 + 1 = 33	or	<input type="checkbox"/> <input type="checkbox"/>	+ <input type="checkbox"/>	= ?

#### SECTION I.—SULPHUROUS ACID.

In the above table, hyposulphurous acid stands first, and, in point of oxygenation, that is its proper place. But it cannot be noticed, until the next article is disposed of, because its formation depends on a salt, composed in part of sulphurous acid, which it behooves us, therefore, to speak of previously.

It will be seen, by the table, that 32, by weight, is the equivalent of sulphurous acid, and that it is composed of equal parts of oxygen and sulphur. Its specific gravity is estimated at 2.222, and one hundred cubic inches weigh 67.776 grains. It is a very pungent and suffocating gas, and may be liquefied by great pressure, or by surrounding it by a freezing mixture of snow and salt. The liquid acid has a specific gravity equal to 1.45, and it boils at 14°, Fahrenheit, producing an intense degree of cold. Water absorbs 33 times its volume of sulphurous acid gas, at the ordinary temperature.



Sulphurous acid may be formed, by burning sulphur in oxygen or common air, but it is more easily prepared, by decomposing sulphuric acid, heating it in a glass retort with some substance which can deprive it of a portion of its oxygen. Almost all vegetable bodies can produce this effect, and a great number of the metals. Mercury is the metal that is generally employed for the preparation of pure sulphurous acid; two parts of it, by weight, being mixed with three of sulphuric acid in a tubulated glass retort, supported on a retort stand, and heated by a chauffer, or spirit lamp. When it is prepared on the small scale, 200 grains of mercury and 300 of sulphuric acid, (about 3 drachms by measure,) may be taken and put into a retort capable of containing about three ounces, (water measure); the beak of the retort should be extremely small.

As water absorbs such a large quantity of this gas, it must be collected in jars over the mercurial trough, which is constructed on the same principles as the water trough, already described. It may be made of wood, marble, or cast iron; the latter is generally preferred, and it is well varnished, to prevent it from rusting. The one in use in the Medical College of Ohio, is made of fine sand-stone. It is about 18 inches long, by 12 wide, and 10 deep. The cavity hollowed out of the stone, is of an oval shape, and about 6 inches deep below the shelf, which is 24 inches wide. A broader stone, placed underneath, and having a deep groove near the edges, serves to catch any mercury that may fall over. One hundred and fifty pounds of quicksilver are sufficient for a trough of this size. Newman\* has cast iron mercurial troughs of two different sizes, the one requiring 65, and the other only 20 pounds of mercury; the large trough allows jars to be used rather more than two inches in diameter, and nine inches long; but for the other, they must not exceed one and a half inches in diameter. Nearly all the most important and interesting experiments, where a mercurial trough is required, may be performed with a small trough, containing only 20 pounds of mercury; and a little practice will soon enable the operator to adjust the beaks of one or two retorts, (by heating and drawing them out at the blowpipe,) to the small jars or bottles, which he must use along with it.

The jars for the mercurial trough must be made at least one-tenth of an inch in thickness, though not more than one or two inches in diameter; they ought also to be ground at the edges, that they may be removed easily on a flat, glass plate, rubbed over with a little gas lute, without losing any of their contents.

\* A philosophical instrument maker in London.

Blotting paper is constantly required, to remove any acid or water that may collect on the surface of the mercury, and when any acid gas has been prepared over it, the mercury should always be washed with water afterwards, and dried with a sponge and blotting paper. A red hot poker, held for a short time in the mercury, enables this to be done more effectually; it is in this manner, also, that the mercury is most conveniently brought to a proper temperature, when it is required to be heated for particular experiments.

The beak of the retort must be placed near the surface of the mercury, that the gas may have to overcome as little resistance as possible, in rising through this heavy fluid; none should be collected till the atmospheric air has been expelled.

The theory of the process is very simple; one equivalent of sulphuric acid, (composed of three of oxygen =  $24 + 16 =$  one equivalent of sulphur,) loses one equivalent of oxygen, 8, which combines with the metallic mercury, and the rest of the oxygen comes away in combination with the sulphur, in the form of sulphurous acid gas; the oxyde of mercury combines with another portion of sulphuric acid, which is not decomposed, and is converted into sulphate of mercury.

To see how readily this gas is absorbed by water, remove one of the jars filled with it, by means of a flat glass plate held firmly to it, or place the thumb or finger on the mouth of a small glass bottle or tube filled with this gas, and Exp. take it off under water; this fluid will instantly combine with it, and be forced up into the tube with explosive violence by the pressure of the atmosphere.

Take a small glass tube, closed at one end, an inch or two long, and about one-third of an inch in diameter; fill it with water, place the thumb upon it, and invert it in the mercurial trough, transferring a small quantity of water from it into another jar, filled with sulphurous acid, in the same manner as gases Exp. are transferred from one jar to another. The mercury will immediately rise in the jar, and the water will be seen to absorb many times its own bulk of gas.

Sulphurous acid cannot support respiration nor combustion; a suspended candle introduced into a jar of this gas, is Exp. immediately extinguished, and a small quantity of it excites coughing, even when largely diluted with air.

Sulphurous acid bleaches a great variety of vegetable coloring matters, sometimes reddening them before the color disappears. The coloring matter is not completely destroyed, and may be often made to reappear by a stronger acid, as the sulphuric, or by an alkali. The most convenient method for showing its

action on the vegetable colors, is to make a solution of sulphurous acid in water, by transmitting it through this fluid as long as it continues to absorb any. The best apparatus, for this purpose, is a series of bottles connected with a retort, in which the gas is generated, and known by the name of Woulfe's apparatus. The gas, escaping from the retort, passes from the first bottle into the second, by a tube which dips under water, and the excess may be conducted into another bottle, as is seen in Fig. 8th, Plate 2d. The water soon becomes highly charged with the acid gas, and may be used for the purpose of exhibiting its bleaching powers.

The solution of sulphurous acid, prepared in this manner, often contains a small portion of sulphuric acid, which must be carefully neutralised by a little potash or soda, (in solution,) adding it drop by drop, till it can destroy the vegetable blues, which it reddens as long as there is any sulphuric acid present.

A very pure solution, of sulphurous acid in water, may be obtained by heating 100 parts of the black oxyde of manganese with 15 of sulphur, in an iron bottle or gun-barrel, placed in the open fire, and transmitting it through water, the sulphur combining with part of the oxygen of the oxyde to form sulphurous acid.

These solutions soon pass into sulphuric acid when exposed to the air, but may be kept for years in stoppered bottles, with little or no change. Sulphurous acid attracts oxygen from a number of metallic oxydes, as gold and mercury, precipitating them in the metallic form. Its salts are all decomposed, with effervescence, by the stronger acids; when exposed to air and moisture, they attract oxygen, like the acid itself, and become sulphates.

The salts formed by the union of sulphurous acid with bases, are called *sulphites*, to distinguish from *sulphates*, which are formed by sulphuric acid.

The sulphurous acid, in form of gas, has been extensively used in medical practice, in what are called the *sulphur vapor baths*. The body, or any particular part, being exposed to the action of this vapor, is soon covered with profuse perspiration; and in this way, principally, is this expedient of any value, as a remedial agent.

The antidotes for the poison of the acid, are the same with those mentioned under the head of nitric acid.

## SECTION II.—SULPHURIC ACID.

This substance was called *oil of vitriol*, in the days of Basil Valentine, who is said to have discovered it in the fifteenth

century. It was called *oil*, because of its oleaginous appearance; and *vitriol*, from the circumstance of its having been obtained from the green vitriol of commerce. This name is still in common use.

The appellation, *sulphuric acid* was given to indicate, with accuracy, its base; and as oxygen was, for a long time, the only agent supposed to be capable of acidifying, the term sulphuric acid, at once pointed to sulphur and oxygen, as its component parts. It contains 16 of the former, (1 equivalent,) and 24 of the latter, (3 equivalents,) giving 40 as the equivalent of *anhydrous* sulphuric acid, that is, acid free of water. It is therefore to be understood, that absolute sulphuric acid never contains a particle of water, and it is therefore called *dry acid*. The equivalent of common sulphuric acid, of the specific gravity 1.845, is 49, and is composed of dry acid 40, (1 equivalent,) and 9 water, (1 equivalent). It boils at  $590^{\circ}$ , and is converted into vapor, which may be readily condensed. This is a stronger acid than that usually found in the shops, which, for the most part, has been weakened, either by the direct addition of water, or by the absorption of moisture from the atmosphere, in consequence of the mouth of the vessel being left unstopped.

Sulphuric acid is seldom or never prepared on the small scale, large quantities of it being manufactured, in this country and elsewhere, as objects of commerce. The processes by which it is obtained, consist in exposing sulphate of iron, (copperas,) to heat in close vessels, when it is distilled over, and collected in a receiver; or in oxygenating sulphurous acid, by the action of nitrous or nitric acid.

In the first process, the water of crystallisation in the crystallised sulphate is, in a great measure, expelled by heating it over the fire, and the acid procured in the subsequent stages of the operation, consists of two equivalents of dry sulphuric acid, 80, combined with one of water. It usually contains a small quantity of sulphurous acid derived from the decomposition of a portion of sulphuric acid, caused partly by the high temperature, and partly by the protoxyde, in the sulphate, attracting an additional quantity of oxygen. In this state, it is known by the name of *glacial oil of vitriol*, or the *fuming sulphuric acid of Nordhausen*, from a manufactory of it at that place, which has now been carried on for a long time.\* Its specific gravity varies from about 1.80 to 1.98; it emits fumes when exposed to the air, makes a hissing noise when dropped into water, and

\* 600 pounds of green vitriol yield 52 pounds of dry concrete acid, if water be entirely excluded.



boils at about  $100^{\circ}$  of F. When exposed to a very gentle heat in a glass retort, connected with a receiver kept cold by ice or snow, nothing comes over but real sulphuric acid, which is condensed in the receiver, and may be obtained in the solid form, by stopping the distillation before any of the water begins to come along with it.

Anhydrous sulphuric acid emits copious fumes on exposure to the air; it is very tough, and converted into common sulphuric acid, by adding the proper quantity of water. It becomes liquid at  $68^{\circ}$ , and boils at a much lower temperature in its pure state, (about  $120^{\circ}$ ,) than when combined with water.

The second process, however, is that which has been most generally adopted for the preparation of this acid. Eight or nine parts of sulphur are mixed with one of nitre, and the mixture burned in a large room or chamber, lined on every side with lead, and covered to the depth of several inches with water. The sulphur is converted into sulphurous acid during its combustion, and a portion of it into sulphuric acid by combining with some of the oxygen of the nitre, nitrous acid and nitric oxide being disengaged; the sulphurous acid combines with the nitrous acid and some watery vapor, forming a crystalline compound which is decomposed by the water at the bottom of the chamber, being converted into sulphuric acid, which remains in combination with the water, and nitric oxide gas, (deutoxide of nitrogen,) which is disengaged. All the nitric oxide, (which is a light gas,) rises in the chamber, and mixing with a fresh quantity of atmospheric air, combines with the oxygen and forms a dense, ruddy vapor, (nitrous acid,) which immediately falls down, in consequence of its great specific gravity; and meeting with more sulphurous acid and watery vapor, a crystalline compound is again formed, which is resolved, as before, into sulphuric acid and nitric oxide. In this manner, a small quantity of nitre may be made to communicate or hand over, as it were, a large quantity of oxygen from the air to the sulphurous acid, and the same series of combinations and decompositions is made to go on, till the water at the bottom of the chamber has become strongly acid. It is then boiled in leaden vessels to expel a part of the water, and the concentration finished in large glass retorts, heated in a sand bath.

The oxygen contained in the nitre, alone, would be inadequate to the complete formation of the acid, which contains a much greater quantity of oxygen than the nitre is capable of affording.

The following diagram exhibits, more distinctly, the nature of the reaction that takes place between the sulphurous and



nitrous acid, when the crystalline compound, which they form with water, is decomposed by a large quantity of this fluid; the part to the left, showing the elementary composition of nitrous acid, &c., and the other, the compounds resulting from its decomposition.

Nitrous acid	{ nit.	14	30	deutoxyde of nitrogen.
	{ oxy.	8		
	{ oxy.	8		
	{ oxy.	8		
	{ oxy.	8		
Sulphurous acid		32	40	sulphuric acid.
Sulphurous acid		32	40	sulphuric acid.
		<hr/> 110	<hr/> 110	

Hence, one equivalent of nitrous acid is sufficient to convert two of sulphurous acid into sulphuric acid, and one of nitric oxyde is disengaged.

In some manufactories, instead of mixing the sulphur with nitre, the nitrous acid which is disengaged during the conversion of sugar into oxalic acid, by nitric acid, is employed to convert the sulphurous into sulphuric acid. The theory of the preparation of sulphuric acid, may be illustrated very beautifully on the small scale, by making sulphurous acid and nitrous acid gases meet together in a glass vessel, that has two necks. I have employed a quart Woulfe's bottle, closing the central aperture, completely. The other apertures are for the reception of the beaks of two small retorts, into one of which, sulphuric acid and mercury are put to form sulphurous acid gas, and into the other, a small quantity of sugar. Heat the first retort with a spirit lamp, until the gas begins to come over, and then pour a few drachms of nitric acid into the retort containing the sugar. Heat the last retort, and nitrous acid fumes will be given out by the decomposition of the nitric acid. If the retorts be large enough to hold six ounces of water when full, 400 grains of mercury, and six drachms of sulphuric acid may be used; and 90 grains of sugar with four drachms of nitric acid, diluted with an equal bulk of water.

When the gases meet in the large bottle, (into which the retorts are fixed by being ground to the tubulures, or having their beaks passed through corks,) a crystalline compound will be deposited on the sides of the vessel, in beautiful dendritical scales, which often cover its whole internal surface. Remove the retorts, when either the sulphurous or nitrous acid ceases to come over, and pour a little water into the bottle; a brisk

effervescence will immediately take place, wherever it comes in contact with the crystalline compound, which is resolved into nitric oxyde and sulphuric acid, the former producing ruddy colored fumes, as it comes in contact with the air. In all these processes, a little watery vapor must be present in the gases, as sulphurous and nitrous acids do not act on each other, when they are perfectly dry. In this experiment, the water, from the diluted acid, which the nitrous acid vapors carry along with them, is sufficient for the purpose.

Pure sulphuric acid, (or hydrosulphuric acid, as it has been termed, from the water which is usually combined with it,) is transparent, colorless, and inodorous, and has a thick, oily appearance, when poured from one vessel to another. It is very acid and corrosive, reddens the vegetable blues, and tastes extremely sour, even when diluted with a very large quantity of water. It absorbs this fluid rapidly from the air,\* and combines with it in all proportions, a considerable elevation of temperature attending the combination.

**Exp.** Pour a few drachms of colored ether into a thin glass tube, of a half inch bore, two or three feet long, and left open at the upper end. Then mix one or two parts of acid with five of water, in a stone vessel, and dip the tube into it. Violent ebullition will speedily follow.

**Exp.** Tie some cotton or flax round a very thin glass vial, having first interposed a few small pieces of phosphorus. Then pour in one part of sulphuric acid and five of water, and the phosphorus will be inflamed. The heat evolved results from condensation.

The degree of condensation may be easily seen, by pouring sulphuric acid into a long glass tube, till it is about half full, filling it gently with water, emptying it into a jug or thin glass flask, and returning it to the tube when cold. The diminution of volume which ensues, marks the degree of condensation which has taken place. One part of water, by weight, mixed with five of acid, causes the temperature to rise from  $50^{\circ}$  to  $300^{\circ}$ , according to Dr Ure. With one part of ice, and the same quantity of acid, the temperature rises to  $212^{\circ}$ ; but with four of ice to one of acid, it falls below zero.

Sulphuric acid has a great affinity for the different salifiable bases, and can disengage almost all the other acids from their combination with them. It forms soluble compounds with the alkalis, but with most of the earths, especially barytes, the

\* Exposed in unstopped bottles, it doubles its weight in a month.

compounds which it forms are very insoluble. Hence, barytes is generally employed as a test of the presence of sulphuric acid, giving a copious white precipitate in all solutions containing it, either in a free state or combined with other substances.

The different sulphates may be formed, by bringing sulphuric acid into contact with their respective bases, great heat being in general produced, so that they must be mixed cautiously together. Some of these are extensively distributed throughout the globe, as the sulphate of lime, and many are prepared in large quantities by artificial operations. The metallic sulphates are decomposed by a red heat, but the earthy and alkaline sulphates resist the action of a much higher temperature; all the latter may be decomposed by heating them along with charcoal, or passing a stream of hydrogen gas over them at a high temperature, the oxygen both of the sulphuric acid and of the metallic oxyde being withdrawn, and a metallic sulphuret remaining. These decompose water, the sulphur combining with the hydrogen and the metal with the oxygen, forming hydrosulphurets, which are dissolved when a sufficient quantity of water is employed.

Sulphuric acid freezes at  $15^{\circ}$ , in the state it is usually met with in commerce, but if diluted with water so as to have a specific gravity of 1.78, it shoots into large crystals; when placed in snow or ice, and will remain in a solid form, if not exposed to a temperature above  $44^{\circ}$ . If the quantity of water is increased, it requires a much lower temperature to cause it to congeal.\*

By passing sulphuric acid, in vapor, through a red hot porcelain tube, in the manner already described under nitric acid, it is resolved into sulphurous acid and oxygen; and if these gases are returned through the tube, they will unite and form sulphuric acid, a fact that will not appear so improbable, when we reflect on the numerous circumstances by which chemical action is influenced. There are many instances, also, where an action of a similar kind occurs. Thus, water may be decomposed by electricity, and its elements again combined by the same agent, so as to form water; if a stream of hydrogen gas be transmitted over the oxyde of iron, at a red heat, it will combine with the oxygen, and water will be produced; and by

\* Diluted with twelve or thirteen per cent. of water, it freezes much sooner than the latter fluid. Neglect or ignorance of this fact has occasioned sad accidents. Mr Parkes, speaking on this subject, says, 'carboy after carboy burst, by the expansion of the acid in the act of freezing and if the remaining carboys had not been immersed in tepid water, not one would have escaped the general wreck.'

transmitting watery vapor over the metallic iron that remains still at a red heat, hydrogen gas may be procured, the oxygen of the water combining with the iron.\*

The sulphuric acid of commerce is never in a state of absolute purity, being always contaminated with a small quantity of sulphate of potash and sulphate of lead; it is the latter that causes the common sulphuric acid to become turbid when diluted with water, sulphate of lead being insoluble in water or diluted sulphuric acid, though the strong acid can dissolve a small quantity. The only method of obtaining the sulphuric acid perfectly pure, is by distillation, a process that requires some precautions in consequence of the small quantity of caloric which becomes latent, when this acid passes to the gaseous state, causing the ebullition to take place suddenly, and producing violent succussions which frequently break the retort. The best method is to fill a retort about one-third full of acid, placing the beak within a long glass tube which is introduced into a receiver; and applying heat by a good charcoal chauffer or spirit lamp. It is not necessary to keep the tube cold by water, as the acid gives out as little heat during its condensation, as it takes up when converted into vapor. To prevent the succussions from taking place, which would almost certainly break the retort, put several pieces of platina wire, or foil, into the retort along with the acid; fragments of glass may be taken when these are not at hand. The quantity of impurities may be ascertained, by boiling a given weight of the sulphuric acid to dryness, in a glass or platina capsule.

There are few agents that are so extensively employed in chemical operations, as sulphuric acid. Sometimes it is used to form important combinations, or to disengage one of the ingredients of a compound by the great affinity which it has to the other; often for the preparation of sulphurous acid; in bleaching and dyeing; in making freezing mixtures; and in communicating oxygen to a variety of substances.

Pure sulphuric acid is colorless, but if a cork be placed in the mouth of a vial of this acid, a little agitation will give it a brown hue, and sometimes a black color. Any carbonaceous matter will have precisely the same action, and hence the bottles containing it, should have glass stoppers. Cancer quacks purposely give a black color to this acid, in order to disguise it, when they wish to use it as a caustic.

When sulphuric acid, or any other concentrated acid, falls on

\* The cases here noticed, are good illustrations of *analysis* and *synthesis*.

clothing, colored, primarily or secondarily, with a blue, red spots are formed. These may be removed, entirely, without injury, if liquid ammonia, or a solution of potash, be quickly applied. The acid is thus neutralised, and the original color is restored.

Sulphuric acid is used in medicine. The *diluted acid* is a better preparation than the *elixir of vitriol*, which, although a diluted acid, too, is weakened, as to its acid powers, by alcohol, and might be called a tincture. A good medicine, not allied to alcohol, is always to be preferred to one necessarily associated with ardent spirits, if its remedial powers are equally great.\*

This acid is used in form of liniment, mixed with sweet oil. It constitutes *Brodie's* liniment.

Its poisonous properties, when taken by the mouth, are to be counteracted by the same agents spoken of when treating of the antidotes for nitric acid.

Occasionally, it does great mischief externally. I knew a lad who had the bad fortune to break a two gallon bottle filled with this acid, and the whole fell on him, from the breast to the feet. He had presence of mind enough to jump into an oil cask, and thus lessened the evil. Notwithstanding this, he was confined for several months with painful ulcers. The expedient he resorted to, was of the best kind, but it was not used sufficiently soon. Oil of almost any kind should be instantly applied, in all such cases.

### SECTION III.—HYPOSULPHUROUS AND HYPOSULPHURIC ACID.

These acids are, comparatively, of much less importance than those which have been already considered; their composition, according to the most recent analyses, has been already stated.

*Hyposulphurous acid* may be formed, by digesting sulphur in a solution of a sulphite, (a compound of sulphurous acid and a salifiable base,) the two equivalents of oxygen in the sulphurous acid combining with an additional quantity of sulphur, and being thereby converted into two equivalents of hyposulphurous acid. It is not easy to procure this acid in a free state, as it is almost immediately decomposed by the reaction of its elements, when detached from the base with which it is combined. Hyposulphurous acid is distinguished, by the peculiar relation

\* I propose to pursue this subject at length, when my engagements will allow me the requisite time.



which it has to the oxyde of silver, combining with it in preference to soda, which is easily separated from this acid by the oxyde; the only instance where a metallic oxyde can separate a fixed alkali from an acid, without the aid of some other affinity. The hyposulphite of soda dissolves chloride of silver, forming a compound which has a very intense sweet taste, not accompanied by any disagreeable astringency, or any thing that could indicate the presence of a metal.

The *hyposulphuric acid* is prepared, by transmitting sulphurous acid through water, in which finely powdered peroxyde of manganese has been suspended; a portion of the oxygen of the oxyde combining with some of the sulphurous acid and forming sulphuric acid, part of which unites with the remaining sulphurous acid, by which the hyposulphuric acid is produced. Both acids remain in combination with oxyde of manganese, and by adding barytes, it is precipitated, the sulphuric acid being also thrown down in combination with part of the barytes, while the hyposulphuric acid unites with the rest, and remains in solution. By cautiously adding sulphuric acid to this solution, the barytes is removed, and the hyposulphuric acid remains in solution. It has not been procured free from water, and is resolved into sulphurous and sulphuric acids when exposed to heat.

Having gone through the compounds formed by the union of sulphur with oxygen, we proceed to notice the products resulting from the combination of sulphur and hydrogen.

#### SECTION IV.—SULPHURETTED HYDROGEN.

Some persons have proposed to call this compound, *hydrosulphuric acid*, because it happens to possess acid properties; but the name placed at the head of this section, appears to me the best we can employ. It indicates, at once, the nature of the compound. The term *hydrosulphuric acid*, is used, by some, for a compound of sulphuric acid and water, and this is a good reason why we should not apply it to the gas under consideration. There are two species of this substance, viz. the *sulphuretted* and *bi-sulphuretted hydrogen*. The equivalent of the first, by weight, is 17, and it is composed of 16 (one equivalent) of sulphur, and 1 (one equivalent) of hydrogen. By volume, its equivalent is one measure. One hundred cubic inches weigh 36.006 grains, and its specific gravity is 1.1805. It is liquefied by a pressure of 17 atmospheres, at 50° Fahrenheit.

Sulphuretted hydrogen is most easily procured by pouring

sulphuric acid, diluted with three or four parts of water, over sulphuret of iron, reduced to small fragments, and collecting the gas that is disengaged, in jars over the pneumatic trough. The materials may be put into a retort, or into a bottle with a bent tube adapted to the tubulure; the apparatus already described and used for the preparation of hydrogen gas, may also be employed here, and when a steady current of gas is required, the flask and bent tube, will be found very convenient. The gas may be made to come over more rapidly by using a stronger acid, and by reducing part of the sulphuret to powder, or by assisting the action by a gentle heat.

Several methods have been proposed for preparing the sulphuret of iron. By exposing a bar of iron to a white heat, in a furnace, or at a smith's forge, and then bringing it in contact with a piece of roll sulphur, it is obtained perhaps more conveniently than in any other way. The iron combines immediately with the sulphur, forming a liquid which speedily becomes a solid brittle mass, with a metallic lustre; it is a protosulphuret of iron, and the great superiority of this process to the others, consists in the protosulphuret, formed in this manner, containing no metallic iron and no excess of sulphur; the former would cause pure hydrogen to be disengaged along with the sulphuretted hydrogen. When more than one equivalent of sulphur is combined with one of iron, it does not afford any gas, when treated with an acid. The roll sulphur should be placed on a stone, and the iron brought in contact with it, gradually bringing them closer and closer together, as the sulphuret is formed; or they may be held over an iron basin filled with water, into which the melted sulphuret falls, taking care always to keep the sulphur and iron in contact with one another. A shower of sparks is thrown off during the action, and if the sulphuret is allowed to fall on the ground, it divides into an infinite number of small globules, presenting a very beautiful appearance; most of it, however, is lost, when it is allowed to fall to the ground in this manner. The iron must be at a white heat, otherwise it merely melts the sulphur, and causes it to take fire, no sulphuret being formed. The roll sulphur should be wrapped round with some cloth or cotton, where it is to be held by the hand, or a glove may be put on to prevent it from falling to pieces. When a long bar of iron cannot be procured, a smaller piece may be held with a pair of pincers; and care must be taken to hold the iron and sulphur in such a way, as to prevent the liquid from running to the hands. If the sulphur is applied to the middle of the bar, when at a proper temperature, it may be divided into two parts, in a few seconds, by pressing

the sulphur gently against it. The iron should be taken out of the forge or furnace, whenever it has become sufficiently hot, as it would be speedily destroyed in the fire at this high temperature.

The protosulphuret of iron may also be obtained, by exposing iron pyrites to heat, in a crucible placed in a furnace, till the excess of sulphur which it contains has been expelled. When neither of these methods can be conveniently followed, it may be procured by mixing four parts of sulphur with seven of iron filings, and exposing the mixture to heat, in a dry Florence flask, (which should not be more than a third full,) resting on the red hot cinders in a good chauffer. In a short time, the sulphur melts and combines with the iron, a rich red glow of light pervading the whole mass during the combination; a cork, with a small piece cut out at the side, (to allow the vapors of sulphur which are formed at first, to be disengaged,) should be put into the neck of the flask, to prevent the free access of the air. The flask need not be kept over the chauffer after the glow of light has begun to appear; and care must be taken not to press the cork too hard in, lest the vapor of the sulphur, not having room to escape freely, should cause an explosion; the flask is broken afterwards, to get the sulphuret. When it is not desired to see the glow of light attending the combination, the mixture may be heated in a covered crucible. The sulphuret prepared in this manner, always has some metallic iron mixed with it, but the sulphuretted hydrogen which it affords, is sufficiently pure for all ordinary experiments.

The theory of the preparation of this gas is very simple. A reaction takes place between every nine parts of water (one equivalent) and 44 (one equivalent) of the sulphuret of iron, the latter being composed of 16 of sulphur and 28 of iron. The oxygen of the water combines with the iron of the sulphuret, forming oxyde of iron, with which one equivalent of the acid unites, converting it into sulphate of iron, which remains in solution, while the sulphur and the hydrogen combine to form sulphuretted hydrogen. The following diagram is intended to illustrate the decomposition.

Before Decomposition.		After Decomposition.	
9 water	{ hyd.	1	17 sulphuretted hydrogen.
	{ oxy.	8	
44 sulphuret	{ sulph.	16	76 sulphate of iron.
of iron	{ iron	28	
40 sulphuric acid		40	
—		—	—
93		93	93
—		—	—

Accordingly, 44 parts of the sulphuret will give 17 of sulphuretted hydrogen, and 40 of real sulphuric acid will be required for the action.

Instead of diluted sulphuric acid, strong muriatic acid may be used with the sulphuret of iron, and when sulphuretted hydrogen is required particularly pure, sulphuret of antimony, reduced to powder, is mixed with five times its weight of muriatic acid, the antimony being oxydated by the water which the acid contains, while the sulphur, combining with the hydrogen that is set at liberty, forms the gas that is evolved. It does not appear, indeed, to be purer than that which is obtained from the sulphuret of iron, as prepared by the first process described, though it is certainly much purer than when the sulphuret prepared from a mixture of sulphur and iron filings, is used. The materials must be heated by a lamp or chauffer, when the sulphuret of antimony is employed.

This gas has also been procured frequently, of late, by the action of diluted acids on the sulphurets of potassium and calcium, prepared by decomposing the sulphates of potash and lime by charcoal, in the manner that will be explained hereafter; the same reaction takes place between the diluted acid and the sulphuret of calcium or potassium, as has been already described, when the sulphuret of iron or antimony is used. Sulphuric acid must not be used with the sulphuret of calcium, as it forms an insoluble compound with lime, (oxyde of calcium,) which would speedily accumulate and prevent the farther action of the acid on the remaining sulphuret; muriatic acid, however, does well, as the muriate of lime is very soluble. Small quantities of sulphuretted hydrogen may be obtained, by subliming sulphur in hydrogen gas.

As sulphuretted hydrogen has not only an extremely offensive odor, (similar to that of sulphureous mineral waters,) but produces severe headache when it is mixed, even in small proportion, with atmospheric air, none should be allowed to escape into the apartment in which it is prepared. A small quantity, indeed, such as is usually lost during its preparation, will not do any harm; but where the materials, from which it has been procured, have been carelessly thrown aside, before the gas has ceased to come, it still goes on accumulating, and often produces very deleterious effects. It is, indeed, much more noxious to animal life, than one would have at first anticipated; a horse dies in air, containing even 1-150th of its bulk of this gas, and dogs and small animals when the quantity is so small as an 1-800dth part; a small bird died, instantly, in air containing a 1-1500dth part. Chaussier states, that many animals die soon,



if they are put into bladders full of this gas, though their heads are left out, and they are allowed to breathe atmospheric air as freely as before.

Sulphuretted hydrogen burns with a pale-blue, lambent flame, when kindled in contact with atmospheric air or oxygen gas. This may be shown, by dipping a lighted candle  
Exp. into a jar of this gas, suspended by means of an iron wire.

During its combustion, the hydrogen combines with the oxygen of the air, forming water, while the sulphur, uniting with the same element, is converted into sulphurous acid. It is evident from this, that one measure of sulphuretted hydrogen will require a measure and a half of oxygen for its complete combustion; for it contains, in a condensed state, its own bulk of hydrogen, which requires half a measure of oxygen to convert it into water, and an equal volume of the vapor of sulphur, with which two equivalents of oxygen (one measure) must be combined, to produce sulphurous acid.

The combustion is never complete in atmospheric air, part of the sulphur being always deposited on the sides of the vessel in which it is inflamed; when mixed with pure oxygen, however, in the proportion just mentioned, the combustion is always complete, and a sharp report accompanies the combination; the gases must, in this case, be mixed in the detonating bottle.

Water absorbssulphuretted hydrogen readily, taking up more than its own volume of this gas when previously deprived of air by boiling, and acquiring its peculiar odor.\* On exposure to the air, part of the sulphuretted hydrogen escapes, and the rest is decomposed, the water acquiring an opalescent appearance, from the deposition of a portion of sulphur, or bi-sulphuretted hydrogen. Sulphuretted hydrogen reddens the infusion of litmus, and possesses the other properties of an acid. The red tint which it communicates, is not permanent, however, from the cause which has just been mentioned; and if the reddened infusion is boiled, it becomes blue immediately, the gas being expelled, as may be easily seen, by heating a small quantity in a Florence flask. Sulphuretted hydrogen-water may be kept for years, in bottles well corked and sealed; they ought also to be filled perfectly full. Perhaps the most convenient method of preparing sulphuretted hydrogen-water is, to fill bottles about two-thirds full with the gas, over the pneumatic trough, and to agitate the remaining water briskly with it, as long as it contin-

\* Several mineral waters in Europe, as well as in this country, contain this gas; some in large and others in small quantities. Its presence is apparent from the very disagreeable smell of the waters.



ues to absorb any, taking out the cork from time to time, to allow atmospheric air to enter, and supply the place of the gas as it combines with the water.

Fuming nitric acid decomposes sulphuretted hydrogen, communicating oxygen to the hydrogen, while the sulphur is deposited; if a piece of thin paper be placed over the bottle containing the gas, whenever the acid is poured in, and the finger pressed very gently upon it, so as to prevent any escaping, the temperature rises so high, that the sulphur and any undecomposed sulphuretted hydrogen immediately take fire, burning with a beautiful flame, and producing a slight detonation; the experiment may be performed with a flask or bottle containing a few cubic inches of the gas, with perfect safety. It succeeds, it has been affirmed, only when the sulphuretted hydrogen has been prepared by diluted sulphuric acid and sulphuret of iron. Exp.

Mix three volumes of sulphuretted hydrogen, with two of sulphurous acid, in a jar over a mercurial trough; they condense into a solid compound, which is the hydrosulphurous acid of Dr Thomson. It has not been applied to any use. Chlorine, iodine, and bromine, decompose this gas, combining with the hydrogen and precipitating the sulphur.\*

Sulphuretted hydrogen is a very important agent in practical chemistry, forming several salts which are often prepared in large quantities in different stages, in the decomposition of sulphates by charcoal, where the salifiable base is required in a pure state, or combined with another acid, as the carbonic, nitric, or muriatic. It is also very much employed, as a reagent, for detecting a number of the metals, or separating them from solutions containing them; as in the detection of arsenic, where the sulphur unites with the metallic arsenic, forming a rich, yellow-colored precipitate, if the liquid under examination contain any. In general, the sulphur combines with the metal in solution, forming a metallic sulphuret which is precipitated, while the oxygen, previously in combination with the metal, goes to the hydrogen of the gas. The sulphuretted hydrogen gas should be transmitted through the liquid to be examined, in Woulfe's apparatus. A bottle with a bent tube adapted to it, (which may be done at the blowpipe,) will also do very well for this purpose, or a small apparatus, similar to Woulfe's, may be easily constructed for the purpose, with some wide-mouthed jars, bottles, or glasses, with corks and bent tubes, fixing the corks

\* It is on this principle, that chloride of lime is useful in destroying the foul odor from privies and other sources.

with a little wax lute or wax cement, if they should not be air tight. A solution of caustic potash should be put into the last glass or bottle, to condense any excess of sulphuretted hydrogen. In lieu of these expedients, I have used a common oil flask, with a glass tube bent like the letter U, only having one leg longer than the other. The materials being placed in the flask, the tube is well secured, and the long leg passed into the solution to be acted on by the gas. The solutions of the common metals which this gas does not decompose, are those of iron, nickel, cobalt, manganese, titanium, and molybdenum. (Dr Henry.)

The most delicate test of sulphuretted hydrogen, is carbonate of lead, which is converted into a sulphuret by the action of this gas, becoming perfectly black. Air containing even a 20,000th part of its bulk of sulphuretted hydrogen, acts on carbonate of lead, the basis of most white paints. Owing to this circumstance, the white paint in privies is very soon tarnished, and becomes almost black. Sulphuretted hydrogen, being largely evolved in such places, has a constant action on the white lead which is the basis of the paint. Any one may satisfy himself of this fact, by filling a bladder with this gas, and forcing it out, so as to come in contact with a perfectly white, painted surface. In an instant, the dark color is produced. It would be good policy, therefore, to use a lead-colored paint, in all such locations.

The hydrosulphurets, (compounds of this acid with the salifiable bases,) may, in general, be formed by passing a stream of sulphuretted hydrogen through a solution of the different bases dissolved or suspended in water, or by indirect processes, where the sulphuretted hydrogen and the base are presented to each other in a nascent state; the most important will be described under their respective bases.

All the hydrosulphurets are decomposed by heat, but the nature of the resulting compound varies according to the base with which the sulphuretted hydrogen is combined. Magnesia parts with all its sulphuretted hydrogen; potash and soda give off hydrogen and sulphuretted hydrogen, part of the sulphur being retained; and the hydrosulphurets of manganese, zinc, iron, tin, and antimony, are converted into water and metallic sulphurets. (Dr Ure.)

Sulphuretted hydrogen gas often induces asphyxia, even to the extent of destroying life. This effect results not only from exposure to privies, but also to drains and sewers.

The first means to be used are, sprinklings with cold vinegar, and frictions with a stiff brush, having previously removed the

patient to the open air. But the chemical treatment seems to be limited pretty much to the action of chlorine, which should be passed to and fro under the nose. If applied immediately, this article is highly serviceable, but it should not be continued too long, for fear of irritating the lungs. The chlorine acts by decomposing the sulphuretted hydrogen, as already noticed. I have somewhere read an account of a physician, who restored a favorite dog, after asphyxia had been induced by this agent, by forcing oxygen gas into his lungs.

#### SECTION V.—BI-SULPHURETTED HYDROGEN.

The prefix *bi*, as explained in our remarks on nomenclature, is an abbreviation of *bis*, twice. The bi-sulphuretted hydrogen contains twice as much sulphur as enters into the sulphuretted hydrogen. Its equivalent, therefore, must be 33; it is composed of 2 equivalents of sulphur = 32, and 1 equivalent of hydrogen = 1. It is prepared, by adding a solution of the sulphuretted hydrosulphuret of potash or lime, to an equal bulk of muriatic acid. These are compounds of bi-sulphuretted hydrogen, and their respective bases, with which the muriatic acid combines, and the bi-sulphuretted hydrogen is slowly deposited in the form of a viscid, oily-looking substance, of a yellow color. Its smell is disagreeable, it is inflammable, and is speedily decomposed, when gently heated or exposed to the air, sulphuretted hydrogen being disengaged, and sulphur deposited. It has not been applied to any use.

The sulphuretted hydrosulphurets may be prepared by digesting solutions of the hydrosulphurets with flowers of sulphur, rubbing the sulphur with a drop or two of the liquid in a mortar, at first, that they may mix more easily together, the sulphur resting on the top of the liquid for a long time, when this precaution is not taken. They may be formed, also, by mixing sulphur with the earths or alkalis, and boiling them in a Florence flask with water; two equivalents of the sulphur combining with one of hydrogen, obtained from an equivalent of water, which is decomposed, while the oxygen of the water probably combines with other two of sulphur, forming hyposulphurous acid; but this has not been satisfactorily ascertained.

The color of these solutions varies, from a greenish yellow to a reddish orange. They are distinguished by their disagreeable odor, bitter taste, and by absorbing oxygen readily from the air, or any gaseous mixture containing it; a property, in consequence of which, they have been much employed, in

eudiometrical experiments. They are all decomposed by the acids, and when the solution is dilute, a copious precipitation of sulphur takes place; and as little or no sulphuretted hydrogen is disengaged, it is likely, that on the addition of an acid which has a great affinity for the base contained in one of these solutions, the oxygen of the hydrosulphurous acid, (supposing that to be the compound which the oxygen of the water that is decomposed forms with the sulphur,) combines with the hydrogen of the bisulphuretted hydrogen, and that the sulphur of both is precipitated.

It is in this manner, that the precipitated sulphur, or *lac sulphuris*, of the London college, is prepared, a sulphuretted hydrosulphuret of lime being formed, by boiling a pound of sulphur with two pounds of lime, in four gallons of water, and adding muriatic acid to the liquid, as soon as it has been filtered through paper. The precipitate must be washed with water till it become tasteless. When sulphuric acid is used, instead of muriatic acid, sulphate of lime is precipitated along with the sulphur, in a very insoluble state; but muriatic acid forms a very soluble salt with lime, and remains in solution.

The *lac sulphuris* was usually prepared with the sulphuric acid, formerly. Sulphate of lime may be easily detected when mixed with sulphur, by exposing a small quantity to heat in a crucible placed in the fire, or still more easily by the blowpipe; the sulphur will be entirely dissipated, but any sulphate of lime, which may be present, will remain.

When these solutions absorb oxygen, they lose their disagreeable smell, their color becomes lighter, and a portion of sulphur is deposited. By long exposure to the air, they become quite colorless, the sulphur being converted into sulphurous and sulphuric acid. They tarnish the metals, and give precipitates with a number of metallic solutions. They must be kept in bottles, well closed with stoppers or corks, made perfectly air tight, with a little wax lute.

---

## CHAPTER V.

### Of Selenium.

This article is of modern discovery, and for the most part, has been obtained in very minute quantities. Berzelius was the first person who met with it; he procured it from sulphur prepared by sublimation from iron pyrites. A few years have



elapsed since it was announced to the scientific world, but it has never been applied to any useful purpose. The thirteenth volume of the *Annals of Philosophy*, contains an account of the process for preparing it; and, recently, it has been advertised for sale, in Germany, at ninety francs per ounce.

It is a solid substance, of a dark brown color, and metallic lustre, melts at  $220^{\circ}$ , boils at  $600^{\circ}$ , forming a yellow colored vapor, which condenses in a powdery form like the flowers of sulphur, but of a red color. It combines with oxygen when heated in the air, producing a strong smell of horse-radish, and forming oxyde of selenium and a small quantity of an acid, at one time called selenic, but now selenious acid, in consequence of another acid compound having been discovered, containing a larger quantity of oxygen than it.

The selenic acid bears a great resemblance, in all its chemical habitudes with the different salifiable bases, to the sulphuric acid; and selenium, indeed, in most of its combinations, produces compounds analogous to those which sulphur forms with the same substances.

Selenium and hydrogen unite when the seleniuret of potassium acts on water, the oxygen of the water combining with the potassium, and forming potash, while the selenium goes to the hydrogen. On adding muriatic acid to the compound of the seleniuretted hydrogen and the potash which remains, the former is disengaged in the gaseous form, while the muriatic acid combines with the potash. It possesses all the properties of an acid.

---

## CHAPTER VI.

### Of Phosphorus.

THIS article was discovered by a German chemist, named Brandt, in the year 1669, while making experiments on urine. Kunckel soon afterwards saw the article, and gave intelligence to a man by the name of Kraft, who purchased the secret from Brandt, for the sum of two hundred dollars. Kunckel, however, prepared it without any knowledge of the process. It was ultimately made familiar to a London apothecary, who made it in such large quantities, that he was able to supply all Europe with it. Hence it was known to chemists, generally, by the name of *English phosphorus*. After this, it was made in France, and its properties were universally known. But no



one, as yet, had employed bones for its formation, and it remained for Scheele to procure it from this source. This discovery was made in 1769, and his process was afterwards improved by Fourcroy and Vanquelin.

Phosphorus is, by common consent, placed among the simple substances; but Sir H. Davy inclined to the opinion that it was a compound, and that hydrogen entered into its composition. But as no decisive experiments have been made to disprove its simple character, we shall continue to give it the usual classification. Its equivalent number is 12; its specific gravity, 1.7. It melts at  $100^{\circ}$ , and boils at  $550^{\circ}$ .

The question has been often asked, whence comes the large quantity of phosphoric matter that is found in human urine and bones? To this quere, a writer in the *Medico-Chirurgical Transactions*, volume tenth, gives the following reply. 'The researches of chemistry have demonstrated, that this matter exists in those animal and vegetable substances, which we use as food. Mr Barry, in experimenting on pharmaceutical extracts, *in vacuo*, found that phosphoric acid was present in all the extracts in a soluble state. And on pursuing his researches, he also found, that the same acid is contained in a vast variety of vegetables, and especially in those which are cultivated for the daily use of man.'

When bones are burnt to whiteness in an open fire, all the animal matter which they contain is destroyed, and nothing remains but a solid mass of a fine white color, consisting, almost entirely, of phosphate of lime, (a compound of phosphoric acid and lime). It is from this that phosphorus is usually prepared; a superphosphate of lime being formed in the first place, by mixing the phosphate in fine powder with three fourths of its weight of sulphuric acid, previously diluted with an equal weight of water. The mixture should be made in a large wedgwood mortar, stirring it constantly, and adding water from time to time, till the mass becomes quite fluid and uniform. The sulphuric acid combines with the greater part of the lime, and the phosphoric acid, set at liberty, attaches itself to a portion of the phosphate which is not decomposed, forming superphosphate of lime, which remains in solution. After the mixture has been kept for a day or two, stirring it frequently, and adding more water to keep it quite thin, it is put into a linen bag. The superphosphate of lime, which is dissolved by the water, filters into a receiver placed below, while the sulphate of lime remains. More water is poured upon it, as long as the liquor which passes through contains much acid, and the mixed solutions are evaporated till they assume a syrupy consistence, when they are to

be mixed with as much powdered charcoal as may render them solid. On exposing this in an earthen retort to a strong heat in a furnace, phosphorus is disengaged, and as it is easily volatilized, it may be collected, by connecting the beak of the retort with a tin tube which is made to dip in water.

In conducting this operation, a very large quantity of gas is disengaged, which inflames when it comes in contact with the air, and the phosphorus congeals, occasionally, in the tin tube, or the neck of the retort, preventing the farther escape of gas. To prevent accident, it should be immediately melted by a chauffer or shovel of red hot cinders, held below the part where the stoppage is supposed to have taken place.

The materials should be well dried by exposure to heat in an iron ladle or crucible, placed over an open fire, before they are put into the retort, to prevent them from swelling and boiling over. The retort should be coated with Willis's lute,\* which prevents a great part of the phosphorus from passing through the pores of the retort, through which it easily forces its way at a high temperature, when this precaution is not adopted. The superphosphate of lime does better for the preparation of phosphorus, than pure phosphoric acid, as it is not so apt to be volatilized.

In this process, the carbon combines with the oxygen of the phosphoric acid, and forms part of the gaseous products which are disengaged; and phosphuretted hydrogen gas is also produced, by the combination of a portion of phosphorus with some of the hydrogen of the water that is always united with this acid, when prepared in the manner that has been described.

If it is required merely to illustrate the principle on which phosphorus is prepared, and not to procure any quantity of it, which may now be always purchased from the manufacturer, thirty or fifty grains of a mixture of phosphoric acid, (or the superphosphate of lime prepared in the manner just mentioned,) with half its weight of charcoal, may be put into a glass tube, sealed at one end, about a foot in length and half an inch in diameter. The tube should be coated with a mixture of two parts of clay and one of sand, previously mixed with cut thread or flax, and then wrapped round with iron wire. The coating need not extend farther than an inch beyond the part to which

\* Made as follows: an ounce of borax is to be dissolved in half a pint of water, and slaked lime added to the solution till a thin paste is obtained. This is to be spread over the retort with a brush, and covered, when dry, with a lute made of linseed oil and slaked lime. It may then be put aside for a day or two to dry slowly, when it will be fit for use.

the mixture reaches when it has been introduced, as this alone is to be exposed to heat. It is placed horizontally in a chauffer, with a hole cut in the side, and a chimney placed over it, to increase the heat; the tube should be gently inclined downwards, to carry off any watery vapor that may be disengaged and condensed on its sides, and the end, which is not coated, had better be drawn out at the blow-pipe when the mixture has been put in, till it is about a quarter or an eighth of an inch in diameter. A green glass tube is much better than one made of flint glass, as it is not so easily melted. A mixture of red hot cinders and charcoal, gives the best fire for this experiment. The phosphorus soon begins to come, condensing along the sides of the tube, and a flame appears at the open end, similar to what is produced by the combustion of phosphorus. If the tube is broken off above the point where it is coated, after the gas ceases to be disengaged, on blowing through it, the phosphorus will immediately take fire and burn with a vivid light.

We have said, that phosphorus was originally prepared from urine, and it is proper to state the process, in general terms. The urine was first evaporated to dryness, and then exposed to a high temperature, in order to enable its carbonaceous matter to decompose the inherent phosphoric salt. Then, on adding a solution of nitrate of lead, phosphate of lead was thrown down, which, on admixture with charcoal, yielded phosphorus by distillation.

If a solution of acetate of lead be added to phosphate of soda, a similar phosphate of lead is precipitated, which, in like manner, may be made to yield phosphorus.

The phosphorus, obtained in these processes, is never pure, having a reddish-brown color, which arises from the presence of some phosphuret of carbon, formed during the distillation. It is purified most effectually by a second distillation; but for ordinary purposes, it will be sufficient to melt it in hot water and press it through chamois leather under water. The distillation of phosphorus may be conducted in a small glass retort, the phosphorus being covered with water to the depth of about a quarter of an inch before applying heat, and the beak of the retort inserted into a receiver containing a small quantity of water, and dipping under its surface. A strong and steady heat is then applied to the bottom of the retort, by an Argand or common spirit lamp; the phosphorus is melted, the water boils and is condensed in the receiver by the cold water, which soon becomes heated to the temperature of  $212^{\circ}$ , and is thereby not so liable to be forced back into the retort when all the water has been distilled over. The heat should then be gradually

increased by bringing the lamp nearer the bottom of the retort, and the phosphorus is speedily volatilised, condensing in the neck of the retort and dropping into the receiver. The vapor of phosphorus remaining in the retort, soon condenses; the water rises in the neck, but as the quantity is so small, atmospheric air immediately enters before it has proceeded far, when it again falls, and continues rising and falling till the retort has become cold. The small quantity of air which enters, in this manner, at a time, prevents any violent action taking place; its oxygen combines with some of the phosphorus still remaining in the retort; a lambent phosphorescent flame appearing in the neck, and at last the retort is filled with nitrogen gas. When cold, the phosphorus adhering to the neck may be removed by boiling more water in it, but this must not be mixed with the rest, as it contains the most of the phosphuret of carbon, which generally condenses in the neck, forming a crust of a beautiful, rich, red color.

I prefer conducting the distillation of phosphorus in this manner, to filling the retort full with water, after the phosphorus to be distilled has been introduced, and replacing it by hydrogen or nitrogen gas.

The phosphorus obtained by distillation, is transparent, and has little or no color. It may be procured in the form of sticks by pouring it, when melted in hot water, into glass tubes slightly tapered and closed at one end by a cork; the tube must also be placed in hot water; the phosphorus being heavier than this liquid, sinks in the tube and displaces the water. It is of a waxy consistence at natural temperatures, and may be easily cut with a knife.

In all experiments with phosphorus, great care must be taken not to allow it to come in contact with air when melted or exposed even to a very gentle heat, if it is not required to kindle it, as it takes fire then, even if slightly pressed between the fingers, or when heated a very little above its melting point. It ought, also, to be cut under water, as it frequently takes fire when cut in the open air, and should never be held by the fingers.\* Accidents are continually taking place, from not attending to these precautions. It must be kept in bottles filled with water, and should not be exposed to the light, as it then soon acquires a crust upon its surface, and becomes of a darker

\* These cautions are necessary for young operators. I am in the habit of taking it in my hands and cutting it in pieces, as I would a stick of candy, but I always keep my hands wet and the phosphorus quite moist.



color. When phosphorus is to be inflamed in atmospheric air or oxygen gas, it must always be well dried on blotting or filtering paper, which prevents it from throwing out sparks.

Phosphorus appears luminous in the dark, when surrounded by atmospheric air, and traces drawn on the wall with it, present a beautiful, phosphorescent appearance; a bason of water should always be at hand when this is done, in case the friction should cause it to take fire. It appears that this luminousness depends upon a small portion of phosphorus being dissolved by the nitrogen of the air, which combines with the oxygen when in this minute state of division; for when phosphorus is enclosed in a jar with pure oxygen, little or no action takes place at ordinary temperatures. When water is boiled with a little phosphorus, part of it rises in vapor along with the steam and renders it luminous when it comes in contact with the air; if the steam is made to issue from a small aperture, it presents a luminous cone which is incapable of inflaming any combustible matter. Phosphorus is insoluble in water, but is dissolved by alcohol, ether, the fixed and volatile\* oils. The solutions are effected easily, by heating these different liquids with the phosphorus in a flask. They are luminous in the dark, when exposed to the air.

When phosphorus is inflamed in atmospheric air, a large quantity of white fumes is formed, which may be collected, by including the phosphorus in a glass jar. These consist of dry phosphoric acid, formed by the combination of the phosphorus with the oxygen of the air, falling down in the jar like flakes of snow. Great heat and light are produced during its combustion.

When phosphorus is burnt in oxygen gas, the combustion is very brilliant, and an intense, dazzling light is produced, which, however, is of short duration. The phosphorus (previously well dried) may be placed in a copper cup suspended by a wire, from a small plate of copper, and kindled by touching it with a hot wire when it has been put into a bottle or vase filled with oxygen. The reader is referred to the chapter on oxygen, for a more particular account of the combustion of phosphorus, in that gas.

Phosphorus not only burns in oxygen gas, but by forcing a stream of that gas upon it, combustion can be made to take place under water. A few very small pieces, not larger than a half

\* There are some exceptions. Dr Ure says, if a few drops of turpentine, rosemary, or lemon be added to a solution of phosphorus in olive oil, the luminous appearance will be lost.



pea, should be put into a glass tumbler, and this should be half filled with boiling water. Then, having a stop-cock bladder filled with the gas, and a tube screwed to the stop-cock, terminating in a fine point, bring the latter into contact with the phosphorus and force the gas directly on it. The combustion is instantly excited, and forms a pleasing experiment. Exp.

In the protoxyde and deutoxyde of nitrogen, phosphorus can also be made to burn, as the reader has observed. We shall find, in the sequel, that it will burn also in chlorine.

Simple friction causes the combustion of phosphorus, as any one may learn by folding a small piece in paper, and rubbing with a piece of wood or any solid substance. Exp.

Dr Bache, of Philadelphia, asserts, that at the temperature of  $60^{\circ}$  F., or upwards, phosphorus is inflamed, simply by coating it with lampblack or fine charcoal, and that this result occurs equally well in the open air and in a close receiver. I have not tried the experiment, but place full confidence in the above statement. Exp.

If phosphorus be put into a bottle with alcohol and well shaken, say for fifteen minutes, it will be reduced to fine particles, resembling very minute crystals. Exp.

The solution of phosphorus in ether, may be made to exhibit the peculiarly luminous character of this substance. If twenty or thirty grains of phosphorus be put into a two ounce vial full of ether and frequently shaken, it will answer, in a few days, for purposes of experiment. If a piece of loaf sugar be soaked with this solution and then dropped into a bowl of hot water in a dark room, luminous coruscations will be seen, which may be increased by agitating the fluid. This phenomenon is supposed to result from a decomposition of the water, and the formation of a gas, shortly to be spoken of, viz. phosphuretted hydrogen. Some allege that the heat of the water simply evaporates the ether, holding in solution a portion of phosphorus in the form of vapor. The phosphorised ether, of which we have been speaking, has been used in medical practice. Exp.

Nothing but phosphoric acid is formed during the combustion of phosphorus in oxygen gas, but when it is heated in air which has been rarefied to a very great degree, or inflamed in a limited quantity of atmospheric air, not sufficient to supply oxygen for its combustion, both phosphorus and phosphoric acids are formed, and these are also mixed with another substance of a red color, the precise composition of which has not been accurately ascertained; it is generally considered as a

compound of phosphorus and oxygen, with a smaller proportion of oxygen than phosphorous acid.

Phosphorus has been employed in medicine, not only in the form of solution in ether, but in substance. The dose, of necessity, has been very small, since a single grain has been known to destroy life. The French have used this article more freely than any other people, but it is not likely to become a popular remedy. It pervades the whole system very rapidly, as is evident from the fact, stated by a French writer, that on opening the abdomen of a female who died after using the article, the entire cavity was luminous, for a short time. Indeed, we see no reason why, under certain circumstances, an evolution of this combustible matter may not occur in the human system, by decomposition of its compounds. May there not be some sort of connexion between such a phenomenon and the tendency to inflammatory action, physical and mental, so obvious in some individuals?

Although phosphorus is poisonous in substance, it is more so when dissolved in oils, ether, &c. In all cases, it gives rise to like symptoms with those which are produced by the mineral acids, and the treatment must be the same. Externally, it causes much mischief. The burns it occasions, are of the most painful kind, if the cuticle be abraded, and they heal with difficulty. If any particles get under the nails, the difficulty is enhanced. When the skin is unbroken, I have found the application of cold water, for hours, the best remedy; the particles of phosphorus, under the nails, should be cleaned out, while the hand is immersed in water, and I am of opinion, that long continued immersion in cold water, will be the best treatment, under ordinary circumstances.

The following table shows the composition of the most important substances, formed by uniting oxygen and hydrogen with phosphorus.

	Oxygen.	Phosp.	
Phosphorous acid	8	+	12 = 20
Phosphoric acid	16	+	12 = 28
Hydrogenated phosphorus	1	+	12 = 13
Bi-hydrogenated do.	2	+	12 = 14

#### SECTION I.—OF PHOSPHOROUS ACID.

Phosphorous acid is a compound of oxygen 8, (one equivalent,) and phosphorus 12, (one equivalent.) It is formed, when phosphorus is inflamed in less air, than is necessary for its rapid com-

bustion, and even when phosphorus is exposed to the air at natural temperatures; mixed, however, in both these cases, with phosphoric acid. The process which is best adapted for preparing it in a pure state, was pointed out by Sir H. Davy. A piece of dry phosphorus is put into a tube retort, and some of the bi-chloride of mercury, in powder, placed over it. On exposing the retort to heat, the phosphorus, as it rises in vapor through the bi-chloride, takes one proportion of chlorine from it, and a limpid fluid condenses in the receiver, viz. the protochloride of phosphorus. On mixing it with water, they mutually decompose each other, the chlorine combining with the hydrogen of a portion of the water, and forming muriatic acid, while the phosphorus takes the oxygen, and is converted into phosphorous acid; by heating the liquid till it becomes of a thick consistence, all the muriatic acid, and most of the water is driven off, and the phosphorous acid still combined with a portion of water, becomes a solid crystalline mass, on cooling.

It has a sour taste and a disagreeable, fetid smell. It reddens the vegetable blues, and when exposed to heat, is decomposed, a portion of phosphorus being disengaged, and phosphoric acid remaining.

To an acid of phosphorus, containing less oxygen than the compound just noticed, the name of *hypophosphorous acid* has been given. It contains two equivalents of phosphorus, 24, to one of oxygen, 8, making 32, for its equivalent. It is prepared by digesting phosphuret of barytes in water, in which process two compounds are formed, viz. phosphate and hypophosphite of barytes. The latter being soluble, is filtered off and decomposed by sulphuric acid; an insoluble sulphate of barytes is precipitated and hypophosphorous acid held in solution. It is an article of very little importance.

#### SECTION II.—OF PHOSPHORIC ACID.

This acid is composed of phosphorus 12, (one equivalent,) and oxygen 16, (two equivalents,) making its equivalent 28. It is prepared most easily, by burning phosphorus in atmospheric air or oxygen, as already described, care being taken to use more phosphorus than can burn freely in it, to prevent the formation of phosphorous acid. It falls down in the form of a large bulky powder, very like snow, and must be immediately put into a stoppered bottle, as it is extremely deliquescent.

This is an expensive process, however, and is never resorted to, except for the purpose of experimental illustration; the acid

is usually procured from the solution of the superphosphate of lime, which always contains a small quantity of the sulphate, when prepared from the phosphate by sulphuric acid. Carbonate of ammonia is added to this liquid till it is completely neutralized, and as long as any precipitation takes place; the excess of phosphoric acid combines with the ammonia, forming phosphate of ammonia, and disengaging carbonic acid with effervescence, while the phosphate of lime, previously retained in solution, by the excess of acid, is precipitated. The sulphate of lime is at the same time decomposed by another portion of the carbonate of ammonia, carbonate of lime being thrown down, and sulphate of ammonia remaining in the solution, along with the phosphate of ammonia. After concentrating the liquid by evaporation, the phosphate of ammonia crystallises, still mixed with the sulphate of ammonia, and on melting these in silver or platina crucibles, the sulphate of ammonia is volatilised, and the ammonia of the phosphate at the same time expelled. The phosphoric acid is melted, forming a transparent and colorless glass as it cools. In this state, Berthier affirms that it still contains one-fourth of its weight of water.

If the solution of the superphosphate of lime is evaporated, as it is procured at first, a solid substance is obtained, similar in appearance to the glacial phosphoric acid procured in the manner which we have just described; it is evident, however, that it must be contaminated with phosphate and sulphate of lime, and if the evaporation has been conducted in glass or earthen vessels, and the dry mass fused in an earthen vessel, it will contain a still greater quantity of impurities, as phosphoric acid acts both on glass and earthen vessels, when its solution in water is concentrated, or when it is fused by exposure to heat. For a great many purposes where phosphoric acid is required, the solution of the superphosphate of lime may be employed, instead of the purified phosphoric acid.

Phosphoric acid may be formed, also, by dropping small pieces of phosphorus through a tube, into a tubulated retort filled half full of nitric acid, and continuing to add it till no more is dissolved; the acid should be heated gently by a lamp or chauffer. The phosphorus takes oxygen from the acid, and a large quantity of nitric oxide is disengaged, which may be collected over the pneumatic trough; when a sufficient quantity has been added, the phosphoric acid may be obtained in the solid form by evaporating the liquid in the retort, and fusing the dry mass afterwards in a crucible.

Phosphoric acid is sublimed by exposure to heat in close vessels, but when combined with a small quantity of water, it



cannot be volatilised. It is very soluble in water and deliquescent, tastes extremely sour, but is not corrosive; it reddens vegetable blues. It is not easily crystallised, and yields phosphorus when heated along with inflammable matter. It acts upon silica and most vessels containing this earth. With the different salifiable bases it forms an important class of salts, the exact composition of which, like that of the other compounds of phosphorus, requires further investigation.

What was formerly called *phosphatic acid*, formed by exposing phosphorus to a moist atmosphere, has now been found to be a mixture of phosphorous and phosphoric acids.

It is well known, that phosphoric acid sometimes exists in the urine, in a free state. That fluid, in infants and nurses, contains very little, because it is called for, by the process of healthy ossification. Richerand affirms, that the elimination of this acid, by the urinary discharge in infants, is the cause of the softness of the bones, that is peculiar to the disease commonly called *rickets*. Paris thinks otherwise, and hints that this disease depends on defective powers of assimilation, whereby the phosphoric acid is not so combined as to be useful, and hence it is voided as excrementitious matter. The urine of old men abounds with this acid, because the work of ossification is completed, and if it were not conveyed from their systems by a proper outlet, it would ossify all the soft parts of the body.

In addition to the compounds which phosphorus forms with oxygen, we are next to notice two that result from its union with hydrogen.

### SECTION III.—PHOSPHURETTED HYDROGEN.

The name of this compound is variously stated in the books. It is called *hydrogenated* phosphorus, *hydroguretted* phosphorus, and *hydruret* of phosphorus. But the name I have selected, is the more usual one, and is equally correct with any of the others. The component parts are, one equivalent of phosphorus, 12, and one of hydrogen, 1, making the whole number 13. It is a gaseous substance, having 0.972 for its specific gravity, and 100 cubic inches weigh 27.5 grains. It was discovered in 1783, by M. Gengembre.

There are two processes for preparing this gas. The first consists in mixing small pieces of phosphorus with water and potash or lime, and exposing the mixture to heat in a glass or metallic retort. With lime, the gas comes away slowly and



steadily, when it begins to be disengaged; but with potash, it is evolved much more rapidly, and more care is required in conducting the process. With a large quantity of water, the gas is long in coming, as the materials do not act upon each other, till the most of it has been expelled. Forty grains of phosphorus, 50 of caustic potash, and 60 drops of water, give this gas very readily, when gently heated in a small retort, (capable of holding an ounce and a half or two ounces, when quite full,) and with very little trouble. The phosphorus should be put in first, in small pieces, that it may be covered by the water, and the potash last, which produces considerable heat as it is dissolved. A spirit lamp, held in the hand, will be found

Exp. most convenient for heating the retort, as the temperature may be easily regulated by holding the lamp at any distance that may be required. A slight explosion generally takes place in the interior of the retort, from the phosphuretted hydrogen, that is first produced, reacting upon the common air which it contains; and as considerable condensation takes place at the same time, if the beak of the retort is placed in the pneumatic trough before this is observed, water will be very apt to be thrown into the retort by the pressure of the atmosphere.

I have used the following method with invariable success. A retort, capable of holding four or five ounces, is filled with a watery solution of caustic potash, and a spirit lamp applied to cause ebullition. Close at hand I have a deep saucer, which, when full, will allow the beak of the retort to be completely immersed. While the solution in the retort (tubulated) is boiling, the pieces of phosphorus are thrown in at the tubulure, and the stopper replaced. Then the beak of the retort is dipped into a saucer, previously filled with hot water, and the action of the spirit lamp is continued. In a few minutes, the gas comes over in abundance. The hot water in the saucer, prevents a fracture of the retort, even though absorption should occur. In the ordinary way, cold water is employed, and on the occurrence of a partial vacuum in the retort, (highly heated,) the water rushes in, and necessarily breaks the vessel.

When a large quantity of this gas is required, it will be found most convenient to prepare it by filling a green glass retort, (a metallic one is still better) with milk of lime\* and chips of phosphorus; 75 grains of phosphorus, with 1500 of slaked lime

\* Made by mixing slaked lime with water, so as to make it of the consistence of cream or milk.

and three or four ounces of water, by measure, will give a sufficient quantity for the experiments which are usually performed with it; the beak of the retort should not be wide, and a tin tube may be attached to it, if it is not sufficiently long. After some time has elapsed, a gas is disengaged, which does not take fire spontaneously; the process may then be discontinued. Exp.

The other method of preparing this gas, is by the action of muriatic acid, diluted with several parts of water, on phosphuret of calcium;\* a purer gas is said to be obtained in this manner, than by any of the preceding processes. For this purpose, a small tubulated retort is filled nearly full with the acid; fragments of the phosphuret are then put in, and the gas collected over the pneumatic trough. Exp. Phosphuret of calcium, indeed, decomposes water without the addition of an acid, and disengages phosphuretted hydrogen gas, but it is procured more easily, by adding the muriatic acid. In both cases, a portion of the water is decomposed, its oxygen combining with the calcium and forming lime, and the hydrogen uniting with the phosphorus.

In the other processes, a portion of the water is also decomposed, its hydrogen combining with part of the phosphorus and forming the gas which is disengaged, while the oxygen converts the rest into phosphoric and hypophosphorous acids, which remain in combination with the potash or lime.

According to Davy, this gas is formed, also, by pouring six parts of sulphuric acid over one of phosphorus, cut into small pieces and mixed with two of granulated zinc and ten of water. I have frequently prepared it in this way.

The horizontal rings which are peculiar to this gas, have not been satisfactorily explained. They are formed, most perfectly, when the gas takes fire in a calm atmosphere. Every bubble of gas first breaks on the surface of the water, then takes fire, and the ring follows.

Place a jar half full of air on the shelf of the pneumatic trough, and allow the gas to rise in it; a vivid flash of light will be seen as before, and all the oxygen will be consumed. Exp.

Place another jar half full of oxygen gas in the same situation; each portion of the gas as it rises in the jar and mixes with the oxygen, produces an extremely vivid flash of light, perhaps as intense, though only momentary, as can be made by artificial means. In performing this experiment, Exp.

\* Made of phosphorus and lime.

care must be taken not to allow the gas to accumulate in the jar without mixing with the oxygen, which frequently takes place from the deposition of a portion of phosphorus on the surface of the water, which rises with the gas, and often enables a large quantity of it to gather into one globule; and whenever this breaks, the jar is thrown down or broken by the violence of the explosion. To prevent gas accumulating in this manner, all that is necessary is to tap the jar with the finger when they begin to appear, or shake it gently over the shelf of the trough on which it is standing.\*

When mixed suddenly with one and a half times its volume of oxygen gas, it is all consumed, and a violent detonation takes place; two strong jars, containing the proper proportions of the gases, must be employed for this purpose, and the one into which the gas is transferred should be more than sufficient to hold both the gases, and wider than the other; it is scarcely necessary to add, that they must be mixed under water, and only a small quantity of the gases employed.

A flash of light appears, also, according to Dr Henry, when this gas is admitted into a flask exhausted as completely of air as can be done by the airpump, showing that there is still a small quantity of air present. To show this, a jar, with a stop-cock fixed to it, is filled with the gas; or it may be only filled in part, if it should be a large jar, and an exhausted flask, which must also be provided with a proper stop-cock, is connected with it. The stop-cock of the jar is opened first, that the phosphuretted hydrogen may have time to act on the small quantity of air between the two stop-cocks, and on opening the other, water is forced from the pneumatic trough into the jar by the pressure of the atmosphere, and the gas which it contained, at the same time, passes through the connector into the exhausted flask.

The products of its combustion, both in air and oxygen, are phosphoric acid and water. Experiments, similar to those which have been described, may be made by transmitting the gas into jars filled with chlorine, a brilliant green light being produced, while it combines with the hydrogen, and also with the phosphorus, forming muriatic acid and perchloride of phosphorus. If a glass retort is taken into a dark room while full of this gas, it presents a very beautiful appearance, as the atmospheric air gradually mixes with the phosphuretted hydrogen;

\* It is affirmed, that the vapor of ether prevents the combustion of phosphuretted hydrogen. This fact furnishes a ready method of guarding against the accidents that often attend the experiments with this gas.

the beaks of the retorts used in all these experiments, should be small, that neither air nor water may enter too quickly.

Phosphuretted hydrogen may be detonated with either the deutoxyde of nitrogen or nitrous oxyde; it is decomposed, also, by iodine and potassium, which combine with the phosphorus. Water absorbs about five per cent., according to Dr Thompson, when it has been previously boiled to expel the air which it usually contains; solutions of the sulphate of copper and chloride of lime absorb it in large quantity, and are employed to ascertain its purity, as they do not absorb any other gases with which it might be mixed. Its smell is very fetid and disagreeable.

The phenomenon called *will-o'-the-wisp*, and some other luminous appearances, both on sea and land, are dependent on the presence of this gas.

#### SECTION IV.—BI-HYDROGENATED PHOSPHORUS.

This is the second compound of hydrogen and phosphorus, and the last which claims our notice, although one or two others have been spoken of. It is composed of two equivalents of hydrogen, 2, and one equivalent of phosphorus, 12, making its whole number 14. As it contains twice as much hydrogen as one equivalent of the gas last examined, I have chosen the name *bi-hydrogenated phosphorus*, as being in strict agreement with fact. Others place the prefix *bi*, before the other ingredient, and say, *bi-phosphuretted hydrogen*. If the composition of the gas, as here given, be correct, the term I have preferred is unexceptionable. This compound always exists in the gaseous form, and is distinguished from the phosphuretted hydrogen, by not taking fire spontaneously, in common air and in oxygen.

This gas is prepared, by exposing the crystalline compound of water and phosphorous acid (formerly noticed) to heat in a small retort, and collecting the gas, that is disengaged, over water. The oxygen of part of the acid and of the water converts the rest of the phosphorous acid into phosphoric acid, and the phosphorus that remains, combines with the hydrogen of the water that is disengaged. The same compound is formed when phosphorous acid and water are exposed to heat, and towards the end of the processes described for the preparation of phosphuretted hydrogen gas; it is also formed when this gas is allowed to stand over water for some time, part of the phosphorus being deposited.

The smell is not so strong and disagreeable as that of the



phosphuretted hydrogen. When mixed with oxygen, it detonates violently on the approach of flame, or when heated to  $300^{\circ}$ . It inflames spontaneously in chlorine, and is absorbed in small quantity by water. In Dr Ure's Dictionary, and in the 31st volume of the *Ann. de Ch. et Phys.*, a detailed account of Dumas's recent investigations with reference to the exact composition of these gases, may be found.

Phosphorus may also be combined with sulphur, and although the compound is not usually named, it might be called a *phosphuret of sulphur*, or a *sulphuret of phosphorus*, at pleasure, if the component parts were joined in equal portions. I have been in the habit of forming a *phosphoric match bottle*, by putting into a small vial a few grains of the flowers of sulphur and two or three small pieces of phosphorus, dried, by laying them in cotton. Let the mixture be shaken and corked, and then set aside for a few days. On opening the cork and quickly passing in a brimstone match, it will be instantly kindled. By care, in closing the bottle immediately after thus using it, the same mixture will serve for several weeks.

## CHAPTER VII.

### Of Carbon.

A REFERENCE to our table of contents will show, that this chapter is intended to embrace the compounds of carbon with oxygen, hydrogen, nitrogen, sulphur, and phosphorus. It will, of course, include many substances, that are usually found under the head of vegetable chemistry, as alcohol, sugar, yeast, ethers, the vegetable acids, &c. These can be studied most profitably, when viewed in their natural relations; and the plan of Dr Reid, in this respect, is greatly preferable to any other that I have seen. It is the one which I shall pursue, extending it, however, so as to include the more important vegetable alkalis.

Carbon, in common with sulphur and phosphorus, is usually called a simple combustible. Its combining, or equivalent number, is 6, by weight, and one measure, by volume. The specific gravity of the vapor of carbon is 0.416, and 100 cubic inches weigh 12.708 grains. Perfectly pure carbon is identically the same with the substance called diamond.

Though carbon has not been obtained in the gaseous form, so as to admit of a series of experiments being made with it in this state, it may be volatilised by the action of a powerful galvanic battery, and it has been customary to represent its equiv-



alent by volume, by a measure equal to that of the equivalent of hydrogen gas. The term carbon is used to signify the pure and inflammable part of charcoal, which, in the manner that it is usually procured, always contains a small proportion of foreign matter.

Charcoal is prepared, on the large scale, by piling wood into cones, which are often of an immense size, covering it, in a great measure, with earth, and setting fire to it through a few air holes, which are closed when it is properly kindled. It is also obtained, in great quantities, by exposing wood to heat in iron cylinders, for the manufacture of gunpowder, where a finer kind of coal, as it is technically termed, is required, a large quantity of inflammable gases, water, tar, and impure acetic or pyroligneous acid being disengaged; the latter is condensed in barrels connected with the cylinders, and purified by subsequent operations. Wood consists of carbon, oxygen, and hydrogen, and in both these processes, the oxygen and hydrogen are driven off by the heat, the different products that are disengaged, being formed by their combining with one another, and with part of the carbon, while the rest of this element remains mixed with any saline matter that may have previously existed in the wood.

To prepare a small quantity of charcoal, a few pieces of wood may be put into a crucible, covered with sand, and exposed to heat in the open fire or in a furnace. When no more gas is disengaged, it may be removed from the fire, but the charcoal must not be taken out till it is cold, to prevent it from taking fire. The wood (a potatoe does equally well) may be cut into the form of a crucible, or in any other shape that the charcoal may be required to have.

Very pure charcoal may be obtained, by exposing several inflammable substances to heat in vessels when they are not brought into contact with the air, as by passing the vapor of alcohol or turpentine through a red hot tube, where a considerable quantity is deposited, in the form of a fine powder. The condensed soot that is obtained from the imperfect combustion of resinous substances, oil, and many other inflammable matters, and known by the name of lampblack, consists principally of charcoal, in a very fine state of division, and when exposed to a red heat for some time, in a covered crucible, to drive off and decompose any volatile matter adhering to it, may be used for the most delicate experiments, where the purest carbon is required.

Charcoal usually retains the form of the wood from which it is prepared. It is black and brittle, insoluble in water, and

has the property of destroying the odor, taste, and color of many substances, for which it is much employed in the arts, and for domestic purposes, as in rendering tainted meat fresh, &c. By filtration through charcoal powder, water that has become putrid, from long keeping in wooden vessels, is made sweet and palatable. By exposing it to a red heat in close vessels, the same quantity of charcoal may be used again and again. On the small scale, it is often used with great advantage in depriving solutions of their color, when they are required to be chemically examined.\* For this purpose, what is termed *animal charcoal* or *ivory black*,† is usually preferred, which is a mixture of charcoal and phosphate of lime, prepared by exposing bones to heat in close vessels, and derived from the decomposition of the gelatine, of which they contain a considerable quantity.

The charcoal obtained from the gas works, or by the smothered combustion of coal, is usually termed coke, and consists of that portion of the carbon which the hydrogen and oxygen have not been able to carry away in the gaseous form, mixed with saline and earthy matter which the coal contains.

Charcoal is infusible by heat, and undergoes little or no change in its appearance or properties, when exposed to heat in close vessels; but when subjected to the action of a powerful galvanic battery, part of it is volatilised, and the lustre of the remainder is increased, while it also becomes so hard as to be able to scratch glass; approaching in no small degree, therefore, in its character, to the diamond, which is composed of pure carbon, and produces the same compound (carbonic acid) during its combustion.

When charcoal has been recently prepared, it has the property of condensing in its pores a large quantity of different gases, without any other changes taking place in their condition or its own. The following table, by Saussure, shows the quantity of different gases which boxwood charcoal absorbs, when allowed to remain in contact with these different gases for twenty-four hours.

Ammonia	90 vols.	Carbonic acid	35 vols.
Muriatic acid	89	Oxygen	9.25
Sulph. hydrogen	65	Nitrogen	7.5
Nitrous oxyde	40	Hydrogen	1.75

\* The rationale of all these effects is wrapt in obscurity, though doubtless dependent on chemical action.

† The well known article *lampblack*, is carbon in a state of light, impalpable powder, and is produced by burning the refuse of pitch and rosin, in peculiar furnaces.

In conducting these experiments, a piece of red hot box-wood charcoal should be plunged under mercury, and allowed to remain so till it is cold, after which it may be introduced into a measured quantity of any of these gases, placed in the same portion of mercury, taking care not to allow it to come in contact with the air. It imbibes about one sixth of its weight of water from the air in a day.

The indestructibility of charcoal, led the ancients to use it in forming their writing ink, which consisted of this substance, in fine powder, mixed with a weak solution of gum arabic.

We are told, that the writings found in Herculaneum, were formed with this sort of ink, and that they are perfectly black. Many ancient works were predicated on the durability of charcoal. Thus the temple of Ephesus was built on wooden piles, which had been charred on the outside, to preserve them. The beams of the greatest works of art at Herculaneum, were changed into charcoal by the lava which overflowed that city; and after the lapse of more than 1700 years, the charcoal has remained entire.

It is further affirmed, that corn, converted into charcoal in the days of Cæsar, yet remains so firm and perfect that the wheat may be distinguished from the rye. It was a custom in those days, to char the outside of all stakes, intended to remain in the ground, especially in damp places, and the practice still prevails.\*

As we saw when on caloric, charcoal is a bad conductor of heat. Mr Parkes has proposed to profit by this peculiarity, in the construction of vessels to be heated by steam. He suggests that these vessels should have treble, instead of double sides, and that the intermediate spaces should be filled with ground charcoal. He thinks, that this arrangement would prevent the escape of heat, insure a given temperature for a great length of time, and thus accomplish a material saving in fuel. That such would be the result, there is no reasonable ground to doubt.

Charcoal is highly inflammable, producing great heat during its combustion; it is much employed as fuel, especially where smoke must be carefully avoided, and it gives no flame when properly prepared.† In oxygen gas, it burns more brilliantly than in atmospheric air, producing vivid sparks; the apparatus,

\* Bishop Watson's Essays.

† This use of charcoal is not free of danger. When burnt in a room unprovided with a chimney, or other suitable vent, it proves destructive to life, by the formation of an irrespirable gas, to be examined presently.

already described for the combustion of charcoal in nitrous oxyde, may be used here. In both cases, the product of the combustion is carbonic acid, the oxygen neither increasing nor diminishing in volume, but becoming heavier by the quantity of carbon which combines with it; every 16 of oxygen take up 6 of carbon.

Even the diamond, the hardest substance in nature, may be inflamed in oxygen gas, and consisting solely of carbon, carbonic acid is the only product of its combustion. It was in this manner that its real nature was ascertained, though Newton conjectured that it was an inflammable substance, long before the experiment was made, from its great refractive power.

Expose some well pounded charcoal to a red heat, in a covered crucible, remove it, and then drop some nitric acid upon it, from a pipette with a long stem. The charcoal takes oxygen from the acid, and a shower of sparks is thrown out.

Exp.

Charcoal, in a minute state of division, decomposes this acid at a much lower temperature.

From the powerful affinity which charcoal has for oxygen at a high temperature, it is constantly employed for deoxidating the metals and a number of other substances. With this element it forms two important compounds, carbonic oxyde and carbonic acid. It unites with hydrogen in several proportions, and with nitrogen it forms cyanogen, the base of prussic acid. It enters into the composition of all the peculiar products of the vegetable and animal kingdom. It is used in the preparation of iron and steel, in the composition of gunpowder, and forms the basis of black paints and printing ink.

Powdered charcoal has long been employed in the practice of medicine. But its principal remedial uses, that seem to be dependent on chemical action, are, to correct the fetor of discharges from the lungs of pulmonary patients, and to arrest the tendency to gangrene, and obviate the disagreeable odor attendant on actual mortification. It is also employed, on the same principle, to remove the peculiar fetor of breath, noticed in some individuals, both by internal administration, and also as a wash to the teeth. For the latter purpose, viz. as a dentifrice, it is, perhaps, the best article in use.

#### SECTION I.—OF CARBONIC OXYDE.

This is a gaseous substance, the discovery of which is, usually, ascribed to Mr Cruikshank, of Woolwich. But, as the subject involves some points of interest in the history of American



chemistry, it is proper to present a brief, but true statement. Professor Woodhouse, formerly of the University of Pennsylvania,\* made a number of experiments with charcoal and scales of iron, which, by their apparent results, shook his faith in the antiphlogistic system. The compound was exposed to a high temperature with a view to determine the nature of the gaseous product. As this was inflammable, he felt inclined to believe, for a moment, that it must be hydrogen, and that, possibly, the doctrine of phlogiston might be founded in fact. M. Adet, a French chemist, proposed to explain the result, by supposing that charcoal contained hydrogen in its pores, and that the inflammable product was thus procured. Dr Priestley rejected these explanations, and triumphed in the additional evidence which he thought was given to his favorite doctrine of phlogiston, by the experiments of Woodhouse. But all were mistaken about the real nature of the product, and, in the heat of controversy, neglected to give the gas a careful examination. To this duty, Mr Cruikshank applied himself in 1801, and while he confirmed the accuracy of Dr Woodhouse's results, relieved his difficulties, by showing that a new gas had been formed by the process, to which he gave the name *gaseous oxyde of carbon*, since called *carbonic oxyde*.†

The equivalent of this gas, by weight, is 14; by volume, one measure. One hundred cubic inches weigh 29.652 grains. Its specific gravity is 0.972.

There are several processes by which this inflammable gas may be procured, most of which consist, essentially, in depriving carbonic acid of half its oxygen, by heating it with some substance which has a great affinity for this element. The best method, perhaps, consists in exposing dried chalk to heat with an equal weight of iron filings, and a small quantity of charcoal, (from a fifth to a tenth part,) in an iron tube retort, and raising its temperature speedily in a good furnace or open fire, till the gas begins to come. The materials should be reduced to as fine a state of division as possible, and the temperature must never be allowed to fall, otherwise the gas soon ceases to come, or carbonic acid is disengaged, instead of carbonic oxyde.

In this process the chalk (which is a carbonate of lime) parts with its carbonic acid on exposure to heat, and the iron and carbon mixed with it, take away one proportion of oxygen, converting it, accordingly, into carbonic oxyde. Either the iron or

\* Died, at Philadelphia, in 1809.

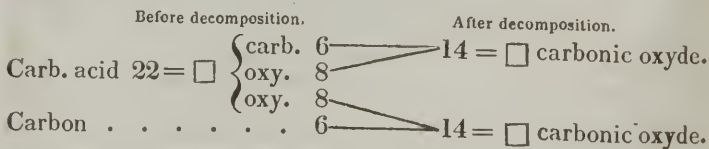
† See Memoirs of the Columbian Chemical Society.



the carbon would do separately, but when they are both taken, there is less risk of the product being contaminated with carbonic acid. A sufficient quantity of gas for showing the general properties of carbonic oxyde may be procured from 250 grains of chalk.

Another process for preparing carbonic oxyde gas, consists in transmitting carbonic acid repeatedly over carbon, at a high temperature, in a porcelain tube. It should be sufficiently long to allow corks to be put in at both ends, without any danger of their being destroyed by the heat when it is made to traverse a furnace, and it should be stuffed with fragments of charcoal; a stop-cock bladder, half full of carbonic acid, being fitted to one of the corks by boring a hole through it, and an empty bladder to the other. When the charcoal has been properly heated, (taking care not to urge the heat too strongly at first, lest the tube should be cracked,) the carbonic acid is passed slowly over the charcoal, from the one bladder to the other, by pressing it gently, and this is repeated several times. The volume of the carbonic acid is doubled, and it is converted entirely into carbonic oxyde. This process, however, is never resorted to except for experimental illustration, as the carbonic oxyde may be procured more easily by the method previously described.

The theory of the action is sufficiently obvious; the carbonic acid is converted into carbonic oxyde, by losing one proportion of oxygen, as is seen by the following diagram; and this combining with part of the carbon, also becomes carbonic oxyde. But carbon occupies the same volume, whether combined with one or two equivalents of oxygen; consequently, a given measure of carbonic acid must double its volume, when converted into carbonic oxyde by carbon.



Carbonic oxyde has been prepared, also, by the decomposition of oxalic acid, which is regarded as a compound of carbonic acid and carbonic oxyde. On exposing it to heat in a small green glass retort, it is resolved into these two gases; but either of the preceding processes is better, as the water of crystallisation which it contains, generally carries a portion of the acid along with it, which is apt to obstruct the beak of the retort and cause an explosion. Dumas proposes to heat the

superoxalate of potash with five times its weight of sulphuric acid, when the latter takes away the potash and water from the oxalic acid, on which it is resolved into carbonic oxyde and carbonic acid. The carbonic acid is separated, by collecting the gases over a trough filled with lime water, the lime combining with it, but not affecting the carbonic oxyde.

When carbonic oxyde is required particularly pure, it ought always to be agitated with lime water, or a very dilute solution of caustic potash, as it generally contains a small quantity of carbonic acid, by whatever process it has been prepared.

Carbonic oxyde is an inflammable gas, transparent and colorless, and burns with a pale-blue and lambent flame. During its combustion, it combines with half its volume of oxygen, and is converted into carbonic acid; a condensation taking place equal to the volume of oxygen consumed. It detonates feebly with pure oxygen, even when mixed in the exact proportions required for converting it into carbonic acid.

Its smell is disagreeable, it is very noxious to animal life, and is sparingly absorbed by water.

## SECTION II.—CARBONIC ACID.

This gaseous substance was discovered by Dr Black, in the year 1757, and was fully examined in his inaugural dissertation. The subject of his thesis was *magnesia alba*, a substance in which he detected the gas now under consideration, and which he called *fixed air*. In addition to magnesia, he found that limestone also yielded this æriform matter, either by the action of heat or the stronger acids; and its formation during respiration, fermentation, and combustion, did not escape the notice of that attentive observer.

Lavoisier was the first person who formed carbonic acid by the direct union of its elements. He burned carbon in oxygen gas, and carbonic acid was the product. A further proof of the nature of this substance, was given by Mr Tennant's analysis. He passed the vapor of phosphorus over carbonate of lime, (chalk,) heated to redness in a glass tube. The phosphorus took oxygen from the carbonic acid of the chalk, and phosphoric acid was formed, while the carbon, thus deprived of its oxygen, was deposited in the form of a light, black powder.

The equivalent of carbonic acid, by weight, is 22, being composed of two equivalents of oxygen, or 16, and one equivalent of carbon, or 6. Its specific gravity is 1.527. One hun-

dred cubic inches weigh 46.597 grains. A pressure equal to 36 atmospheres, renders it liquid.

Carbonic acid is most easily procured, by pouring muriatic acid, diluted with an equal bulk of water, on fragments of white marble, a compound of carbonic acid and lime. The muriatic acid combines with the latter, and the carbonic acid is disengaged with effervescence. The materials may be put into a retort, or a bottle or flask with a bent tube, (Fig. 12th, Plate 2d,) adapted to it, may be employed, as no heat is required, and the gas collected in jars over the pneumatic trough. Two ounces of muriatic acid, by measure, may be taken for every 300 grains of marble; and, as every 100 grains of marble contain 44 grains of carbonic acid, which occupy very nearly the space of 100 cubic inches at ordinary temperatures, it is easy to calculate the quantity of marble required to afford a given bulk of gas.

Carbonic acid is a transparent and colorless gas, heavier than air, and incapable of supporting combustion or respiration. Animals are killed immediately when confined in it, and fatal accidents are frequently occurring in old wells and pits, at the bottom of which it often accumulates in considerable quantity. A burning candle, therefore, should always be let down before any person descends; it will be extinguished if the air is loaded with this gas, and incapable of supporting respiration.\*

When charcoal is burned in atmospheric air, carbonic acid is formed; and if this process be allowed in an unventilated chamber, the air will be so completely surcharged with the carbonic acid, as to be unfit for respiration. Death is frequently the result of such mistakes.

From its great specific gravity, this gas may be poured from one jar into another in the same manner as water, and a jar left full of it with the cover off and the mouth turned upwards, will retain it for a considerable time. A candle may be extinguished, also, by pouring the gas upon it, and the experiment may be varied in several ways.

Pure carbonic acid does not pass into the lungs when an attempt is made to inspire it; a spasm of the glottis, which immediately takes place, prevents its entrance; but if it is diluted with a considerable quantity of air, it then passes freely into the lungs, and though it is scarcely perceptible at first, it soon gains upon the system, and has often proved fatal in a close apartment, where a charcoal chauffer has been used, as we have had occasion to notice already.

\* The usual practice of passing down a tub or bucket, containing freshly slaked lime, is a good one. The lime absorbs the carbonic acid, and it is thus abstracted from the air.

Water absorbs its own volume of carbonic acid gas, or rather more, and, by pressure, it may be made to take up a much larger quantity. A solution is easily made, by filling a bottle about two-thirds full of the gas over the pneumatic trough, and shaking the remaining water with it as long as it continues to absorb any. It may be kept for years, if it is put into bottles which are filled quite full and properly secured.

Carbonic acid water is sometimes prepared, on the small scale, by passing the gas through water, in the bottles of Woulfe's apparatus. Another apparatus, invented by Dr Nooth, is also occasionally used for this purpose. It consists of three glass vessels, which are made to fit accurately to each other by grinding. In the lower vessel, the carbonic acid is prepared from a mixture of equal weights of sulphuric acid and chalk, the acid being previously diluted with twelve times its weight of water, and the apparatus inclined to one side, by a small piece of wood placed below it, that all the chalk (which should be finely powdered) may not mix with it at once. The middle vessel is filled with the water to be impregnated with the gas, a valve at the bottom allowing the carbonic acid to arise into it from the lower vessel, when it is accumulated in sufficient quantity to overcome the pressure of the liquid above, but constructed so as to prevent any of the solution from falling into the lower vessel. When the gas accumulates in the middle vessel, it presses upon the surface of the liquid, a part of which is forced into the vessel above, and when the whole water has been completely saturated, the excess of carbonic acid, after passing into the upper vessel, lifts up the conical stopper at the top, which returns to its place the moment it has escaped. Before using it, care must be had that the valve moves readily, on blowing through it.\*

Carbonic acid water has a pleasant, acidulous taste, sparkles when poured from one vessel to another, but loses its agreeable pungency when exposed to heat, or to the air for some time, from the acid gas escaping. This also takes place when it is frozen, or placed in the exhausted receiver of an airpump.

Pour a little of the infusion of litmus into some carbonic acid water, it will be immediately reddened; but the original color will be restored by exposing it to heat. Exp.

Mix another portion with some lime water in a glass vessel; it will immediately become milky, carbonate of lime being formed, which is insoluble. Exp. Add an additional quantity of the acid water, and the liquid will become clear,

\* The water thus prepared, is the same with the artificial Seltzer water.



as a supercarbonate of lime is then formed, which is soluble. If it is now exposed to heat, the excess of carbonic acid will be disengaged, and it will again become turbid.

Pour some lime water on a flat plate, and expose it to the air; a pellicle will speedily form on its surface, from the carbonic acid in the air combining with the lime.

Breathe through some lime water in a bottle, by a bent tube. The carbonic acid, produced during respiration, will combine with the lime, and render the liquid turbid.

Fill a tube, about 12 or 18 inches long, with carbonic acid gas, over the pneumatic trough; close it with the finger or thumb, and place it in a small cup of mercury, or on the shelf of a mercurial trough. Introduce a small quantity of a strong solution of caustic potash, taking care not to allow any air to enter. If the carbonic acid is quite pure, the potash will absorb it all, carbonate of potash being formed, and the mercury will rise in the tube; if any common air was mixed with it, the bulk of the remaining gas will indicate the quantity, and deducting this from the total amount of gas, we have the quantity of real carbonic acid which it contains.

Carbonic acid combines with the different salifiable bases, forming a very well defined class of salts. It has, however, but a feeble attraction for them, compared with almost all other acids, and may be easily displaced, assuming at the same time the gaseous form, when it is not retained in solution by a large quantity of water. The carbonates accordingly effervesce with most other acids, and by noting the quantity of carbonic acid disengaged in this manner, the quantity of carbonates, in any saline mass, may be estimated. This may be done, either by measuring the volume of carbonic acid, which a given weight of mixed salts affords, or by ascertaining its weight.

In the first case, the easiest method of proceeding is to fill a long tube (closed at one end, and capable of containing two or three cubic inches,) nearly full of mercury, filling it completely, afterwards, with muriatic acid, diluted with an equal quantity of water. The thumb is placed over this, after dipping it in oil, or rubbing it over with a little gas lute, the tube inverted and placed in a cup of mercury. One or two grains of the solid salt are then introduced into the tube, (the experiment is most easily performed with a fragment of some carbonate,) and the moment it rises to the top, and comes in contact with the acid, the carbonic acid is disengaged with effervescence, depressing the mercury, and its amount is estimated by examining the volume which it occupies, and making the usual corrections; one equivalent of carbonic acid indicating one equivalent of a carbonate, whatever may be the nature of the base.



In the other method, which is more generally adopted, a thin glass flask or bottle, is placed on one of the scales of a balance, with some muriatic acid, and accurately counterpoised along with a given weight of the substance under examination, and a long tube, twice bent upon itself, passing through a cork, which fits to the mouth of the flask. This tube is put in when the acid and carbonates are mixed together, to prevent any loss from particles of liquid that are thrown up during the effervescence; and it is evident, that, by adding weights to the scale on which the glass vessel is placed, (when the effervescence has finished,) till it is again counterpoised, they will indicate the quantity of carbonic acid that has been evolved, making allowance for what may still remain within the apparatus.

Carbonic acid exists abundantly in nature,\* especially in combination with lime, forming 44-100dths of limestone, and other kinds of carbonate of lime. In the Grotto del Cano,† and other parts of the globe, it flows out in a continued stream, in the gaseous form; it constitutes the characteristic ingredient of the carbonated mineral waters, and is formed, in large quantities, by the respiration of animals, during combustion and fermentation, and by those changes which dead vegetable and animal matter are continually undergoing at the surface of the earth. It is generally supposed that carbonic acid is prevented from accumulating, and the purity of the atmosphere maintained, by the vegetable world, the leaves of plants decomposing carbonic acid in the sunshine and exhaling oxygen; in the shade, however, they absorb oxygen, and produce carbonic acid.

Although carbonic acid gas cannot be taken into the lungs, but on the contrary, is thrown off as excrementitious, so soon as formed there, yet it may be received into the stomach, freely, and with happy results.

It not only gives life to the malt liquors which are so largely used, but also enters into most of the wines, and into nearly all the artificial mineral waters. If a bottle of porter remain un-

\* The miners call it *choke damp* or *choke vapor*.

† This natural curiosity has a descent of two or three small steps, by which the curious arrive at the bottom. A man may enter with safety, and remain for some considerable time, because the gas, by reason of its specific gravity, falls towards the bottom of the cave. But a dog or other animal, whose head will not be on a level with the surface of earth, beyond the mouth of the Grotto, when forced in, very soon staggers and falls. A man resides in the vicinity, and waits on travellers, who desire to see this work of nature, and he is provided with dogs to amuse his company. He keeps them in the cave long enough to show the deleterious effects of the gas, and restores them, when drawn out, by repeated affusions of cold water.

corked for a short time, it becomes flat, owing to the escape of this gas. The insipidity of water that has been boiled and then cooled, depends on the dissipation of the carbonic acid, which it originally contained, by the agency of heat. We shall recur to this subject, when we take up the analysis of mineral waters.

In the use of the potations, which are known to physicians as *effervescing draughts*, much carbonic acid gas is evolved in the stomach, and not a little of the advantage resulting from these medicines, depends on the action of this gas. It is carried into the stomach also in various articles of food, and is very often formed in that organ, by chemical agency.

Physicians, moreover, employ carbonic acid gas, externally, as in the yeast poultice, the efficacy of which is owing, chiefly, to the constant escape of this æriform matter.

But we have alluded to the fatal effects sometimes induced by this gas, when taken into the lungs. It causes the disease called *asphyxia*. Although much has been written about the proper treatment of persons thus affected, I know of no remedy, that can be said to act, purely on chemical principles; none that, by decomposition, can merit the name of *antidote*. Removal to the open, free, pure air, affusions of cold water, external irritants, and the inflation of the lungs, with oxygen gas, either pure or diluted, seem to make up the entire list of means that good sense or sound philosophy direct us to employ.

### SECTION III.—OF CARBURETTED HYDROGEN OR OLEFIANT GAS.

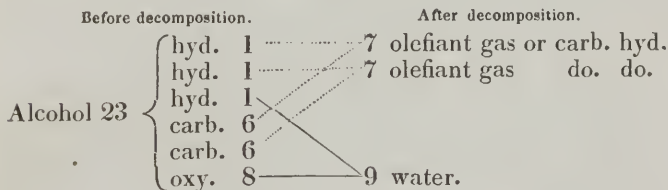
Chemists have differed in their appellations of this gas. It has been called *bi-carburetted hydrogen*, *percarburetted hydrogen*, *hydroguret of carbon*, *hydruret of carbon*, &c. The term *olefiant gas*, was given by the Dutch chemists who discovered it, on account of its forming an oily-like liquid with chlorine. The discovery was made in 1796.

It is a compound of one equivalent of carbon, 6, and one equivalent of hydrogen, 1, making in all, 7, by weight; while, by volume, it is a half measure. Those who call its volume a *whole* measure, make the equivalent, by weight, 14. Now, as the gas is made up of one equivalent of each of its constituents, it might be called *hydrogenated carbon*, as well as *carburetted hydrogen*; but as the carbon has the preponderance over the hydrogen, by weight, in the ratio of 6 to 1, I have thought proper to employ the term which is placed at the head of this article. The specific gravity of the gas before us, is 0.972, and 100 cubic inches weigh 29.652 grains. As it is considerably

heavier than the gas which we shall next examine, it has sometimes been denominated *heavy* carburetted hydrogen gas.

This gas is most easily prepared, by mixing one part, by measure, of alcohol with three of strong sulphuric acid, in a glass retort, and exposing the mixture to a gentle heat. The retort should not be filled more than a third full, and when only a small quantity of the gas is required, half an ounce of alcohol, with a proper quantity of sulphuric acid, will be found quite sufficient. The alcohol and the acid must be shaken together before the heat is applied. A little ether is formed at first, and towards the end of the process, sulphurous acid, carbonic oxide, and bi-hydrogenated carbon, (another compound of carbon and hydrogen,) are disengaged; the mixture also becomes quite black from the deposition of carbon, and is very apt to boil over.

To understand the nature of the changes that take place in this process, it must be recollected, that every 23 parts or one equivalent of alcohol, is composed of three parts of hydrogen, (three equivalents,) twelve of carbon, (two equivalents,) and eight of oxygen, (one equivalent); so that it may be regarded as a compound of one equivalent of water and two of olefiant gas, for the different elements are present, exactly in the proportions necessary to form these compounds. The water then may be said to combine with the sulphuric acid, while the olefiant gas is disengaged; the new arrangement which the elements of the alcohol assume, is represented in the following diagram.



The two proportions of olefiant gas which are disengaged from each equivalent of alcohol at the commencement of the process, often combine with an equivalent of alcohol which is not decomposed, forming one equivalent of ether, which explains its appearance; the black color which the liquid assumes afterwards, and the formation of bi-hydrogenated carbon, arise from the elements of the alcohol arranging themselves in a different manner, which will be readily understood from the annexed diagram.

Before decomposition.		After decomposition.	
Alcohol 23	hyd. 1	8	bi-hydrogenated carbon.
	hyd. 1		
	hyd. 1		
	carb. 6	6	carbon precipitated.
	carb. 6		
	oxy. 8		
		9	water.

But no sooner is the carbon precipitated, than it begins to react upon the sulphuric acid, taking one equivalent of oxygen from it, and being converted into carbonic oxyde, while the sulphuric acid becomes sulphurous acid, in consequence of losing this proportion of oxygen; a small quantity of carbonic acid is also formed towards the end of the process. The proportion of these gases that is disengaged in the succeeding stages of the process, becomes greater and greater as it proceeds; at first, the olefiant gas is very pure, and the sulphurous and carbonic acids, formed afterwards, may be removed by caustic potash.

A large quantity of this gas, mixed with other inflammable compounds of carbon and hydrogen, may be procured by exposing oil\* and resinous or fatty substances to a red heat, in close vessels. It is in this manner that what is termed oil gas, is prepared on a large scale; it owes its great illuminating power, principally, to the olefiant gas which it contains. The process may be imitated on the small scale, in the gun-barrel apparatus, for making hydrogen, using oil instead of water, and putting in iron shavings or fragments of earthen ware, merely to extend the surface. The tube must be brought to a good red heat, before the oil is allowed to drop into it. All these inflammable substances are composed, almost entirely, of carbon and hydrogen, and, by exposure to a high temperature, their elements are made to arrange themselves, so as to form a large quantity of gaseous matter, while a considerable portion of carbon is deposited.

An inflammable gas may be obtained from alcohol, in the same way, but it does not burn with such a rich flame, and contains carbonic oxyde and bi-hydrogenated carbon, along with olefiant gas.

The common bituminous coal, also, yields an inflammable

\* By a recent notice in one of the Philadelphia newspapers, we observe, that some one has obtained a patent for a lamp to hang in halls and other places, in which the oil is changed into an illuminating gas, by a process which goes on within the body of the lamp. We have not yet seen a particular description.

gas, containing a large amount of olefiant gas. It is obtained by heating pieces of the coal, in an iron bottle, placed in a furnace; tar, water, and some carbonate of ammonia pass over, at the same time.

The gas that is disengaged from coal, does not consist solely of compounds of carbon and hydrogen, but contains also carbonic oxyde, carbonic acid, nitrogen, and sulphuretted hydrogen; the proportion of these is not very great, though sufficient to render the gas less luminous during its combustion, and impart other sensible properties to it. The ammonia is produced by the combination of part of the nitrogen and hydrogen of the coal, when the other products are formed, and part of the carbonic acid unites with it, and converts it into the carbonate of ammonia.

In several of the large cities in the United States, as well as in Europe, the streets, manufactories, public buildings, &c., are illuminated with this gas. The coal is placed in large iron retorts, and exposed to a proper degree of heat. A large portion of tar is formed in the process, which is received in an intermediate vessel, and if the gas is required to be pure, it is passed through lime water, to free it from carbonic acid and other gaseous matters. The pure gas is accumulated in a mammoth receiver, or gas-holder, connected with which are pipes for conveying it in all directions.

Olefiant gas burns with a rich, white flame when lighted with a match in contact with the air, and consumes a large quantity of oxygen, requiring three times its bulk of this gas for its complete combustion. With pure oxygen it detonates, violently, when the gases are mixed in the above proportion; a strong detonating bottle must be used in performing this experiment, and it should be wrapped round with a towel, in case of accident. The products of its combustion are water and carbonic acid. In the following diagram the theory of the action is represented, as also the portion of oxygen required, both by weight and by measure, with the quantities of the resulting products.

Before decomposition.		After decomposition.	
7 $\square$ olefiant gas =	{ carb. $\square$ 6 hyd. $\square$ 1	22 $\square$ carbonic acid.	
24 $\square$ oxygen =	{ oxy. $\square$ 8 oxy. $\square$ 8 oxy. $\square$ 8	9 $\square$ water.	
—		—	
31		31	
—		—	



As the watery vapor is immediately condensed at natural temperatures, one measure of carbonic acid is the only gaseous product that results from the detonation of the half measure of olefiant gas.

Water absorbs about an eighth of its volume of olefiant gas, but does not acquire any particular properties. By passing electric sparks through this gas, or transmitting it through a red hot porcelain tube, it is decomposed, the hydrogen resuming its original volume, while the charcoal is deposited. It can combine with chlorine, iodine, and bromine. Some more experiments, which may be performed easily with olefiant gas, will be described under chlorine.

I have been in the practice of exhibiting the illuminating powers of this gas to my class, in the following manner. A large stop-cock receiver is filled with the pure gas, formed by the action of heat on sulphuric acid and alcohol. To the stop-cock of this receiver, a tube, furnished with several gas burners, is screwed on, through which the gas will pass freely, on turning the stop-cock. The receiver is to be gently moved off the shelf and depressed into the water of the pneumatic cistern, and at the same moment, the stop-cock is turned, and a lighted taper applied to the apertures in the gas burners. By gradually depressing the receiver, the gas is regularly forced out, and will burn beautifully in a variety of figures, until the whole quantity is consumed.

#### SECTION IV.—OF BI-HYDROGENATED CARBON OR LIGHT CARBURETTED HYDROGEN.

This gas been called the *inflammable air of marshes*, *hydrocarburet*, and *protocarburet of hydrogen*; it is also named *subcarburetted hydrogen*; and by the miners, it is known as the *fire damp*, or inflammable air that causes the explosions in coal mines.

It is composed, by weight, of 6 carbon, (one equivalent,) and 2 hydrogen, (two equivalents,) making the combining number, 8; by volume, it is a whole measure. Its specific gravity is 0.555, and 100 cubic inches weigh 16.914 grains, so that it is much lighter than the *olefiant gas*, already considered. It is procured most easily, by stirring the mud at the bottom of stagnant pools, and collecting the gas, which is disengaged, in glass vessels inverted over and full of water, a large quantity being formed there by the decomposition of dead vegetable matter. No process has been pointed out, by which it may be obtained with facility in a pure state. It is almost always mixed with a

small quantity of carbonic acid, when obtained in the manner we have described, and also when procured from a coal mine; by washing it with lime or potash water, the acid gas is removed.

Small quantities are formed in a great number of processes, as in some of those for the preparation of olefiant gas; and most vegetable substances afford a considerable quantity, when they are exposed to heat in close vessels.

The bi-hydrogenated carbon has neither taste nor smell when pure, cannot support combustion nor respiration, and is absorbed very sparingly by water. It burns with a yellow flame, and consumes twice its volume of oxygen during combustion; the two equivalents of hydrogen which it contains in a condensed state, combining with two equivalents of oxygen, and forming two of water, while the carbon unites with as much more oxygen, and is converted into carbonic acid.

The coal gas, made in the usual way, consists of olefiant gas and bi-hydrogenated carbon, and may be employed for showing the operation of Sir Humphrey Davy's safety-lamp, as a sufficient quantity of bi-hydrogenated carbon cannot be so easily procured, in a separate state. Mix the gas with various proportions of atmospheric air in a strong jar or detonating bottle. The pure bi-hydrogenated carbon does not detonate, unless mixed with more than four times its volume of atmospheric air, and the explosion is feeble, till seven or eight times its bulk is added. With more than fourteen times its volume of air, it does not form an explosive mixture, a candle merely burning in it with an enlarged flame. The proportion of air for performing the same experiments with the coal gas must be greater, since, by reason of the olefiant gas which it contains, more oxygen is required for its combustion.

Mix it in the same manner with different proportions of pure oxygen; the detonation takes place, violently, with two or three times its bulk of this gas. The products of the combustion are water and carbonic acid. Exp.

Introduce the safety-lamp, when burning, into mixtures of air and coal gas, in various proportions; the mixture passes through the meshes of the wire gauze and burns within it, the flame extending and striking against the top of the gauze, when it is again reflected downwards, but never inflaming the inflammable mixture without the cage. Exp.

Connect a flexible tube with the gasometer, and direct a large stream of the gas upon the lamp held in the open air, after it has been lighted and properly adjusted. By varying the distance at which the lamp is held from the gas, it will either burn with a larger flame, or be completely extinguished,

when the quantity of gas which falls upon it is so great as to prevent the admission of atmospheric air; the inflammable atmosphere around, however, is never inflamed.

The construction of the safety-lamp cannot be well understood, without an examination of the instrument, although it is a very simple contrivance. A wire recurved at the top passes through the cistern containing the oil, and without the lamp; it is called the safety trimmer, as with it the wick of the lamp may be trimmed, without taking off the gauze. The wire gauze is fixed to a brass rim which is made to screw upon the cistern containing the oil. It admits the air, to support the combustion, through its apertures, and has a double top, one being placed at some distance from the other, as an additional security.

The cause why the burning wick of the safety-lamp does not fire an inflammable mixture, exterior to the gauze, though it can enter freely and burns within, has been satisfactorily explained. A very great heat is required to inflame most bodies in the gaseous state, and when they become luminous, so as to constitute flame, their temperature is above the white heat of solid bodies. But when any gaseous matter in the state of flame passes through the minute apertures of the wire gauze, its temperature is so much reduced, that it ceases to be luminous, and is incapable of inflaming any explosive mixture, that may be without. The metal, also, of which it is composed, being a good conductor of caloric, gives it off quickly to the external air, when heated, so that it is always sufficiently cool to reduce the temperature of any gas that may be burning within the gauze, as it passes through, nor does it ever become so hot, as to be able to inflame an explosive mixture itself.

The safety-lamp is justly ranked among the most brilliant of Sir Humphrey Davy's discoveries. Hundreds of lives were lost by the sudden accumulation and explosion of the gas in mines, before this instrument was invented. Now, the miners work with perfect safety, and an accident is seldom reported.

Several other compounds of carbon and hydrogen have been announced, within a few years past. Two of them have been pointed out by Mr Faraday, and are procured from the liquid which is collected during the condensation of oil gas, at the establishments for making this article.

One of these has been called *quadro-carburetted hydrogen gas*, and is composed of the same relative proportions of carbon and hydrogen as are found in olefiant gas, only that two equivalents of each are present. It is obtained, simply, by heating the above named liquid in a small retort, with the hand, and condensing the product of the distillation in a receiver, cooled to

zero by a freezing mixture. The liquid obtained, boils at a temperature below the freezing point of water, forming a vapor which is highly inflammable, and burns with a very dense flame. The products of its combustion are water and carbonic acid, and it requires exactly twice as much oxygen for combustion as olefiant gas, consisting of two equivalents of hydrogen, and two of carbon, condensed into half a measure. The fact that the same elements, united together exactly in the same relative proportion, could still combine so as to form compounds with different properties, is certainly one that could not have been anticipated; there is nothing paradoxical in it, however, as the elementary particles may be associated in a different manner in each. Mr Dalton discovered a compound of carbon and hydrogen in oil gas, in which three equivalents of each of these substances are condensed into one volume. It has been called *superolefiant gas*.

*Naphtha*, an oil which is obtained by distillation from coal tar, similar to that which is found native in many parts of the globe, and known by the name of mineral naphtha, is composed, also, of carbon and hydrogen, united in the same proportions as in the preceding compounds, but liquid at natural temperatures, and producing a dense vapor when volatilised, one measure of which contains, in a condensed state, six measures of carbon vapor, and six of hydrogen; it requires nine times its bulk of oxygen for combustion.

This liquid, containing no oxygen, is used for preserving potassium,\* a substance that takes oxygen rapidly from all fluids in which this element exists, and from atmospheric air. It is purified by mixing it with a small quantity of sulphuric acid, and redistilling it in a glass retort. It does not combine with water, but unites readily with sulphuric ether, alcohol, and oils. It is dissolved by some of the gases, and it is supposed to impart to coal gas, the offensive smell which it always has. Caoutchouc is dissolved by naphtha, and this fluid is used to make the solution of this substance, employed in the manufacture of waterproof cloth.

The *bi-carburet of hydrogen* is another compound that was discovered by Mr Faraday, and prepared from the same liquid as the other, by condensing and repeatedly distilling the product obtained on heating it to a temperature of about  $180^{\circ}$ . It is a colorless and transparent liquid at ordinary temperatures,

\* I am inclined to believe, from some little experience, that none of the substitutes for naphtha, that have been suggested, will answer the purpose. In this opinion, I find myself supported by the testimony of Dr Gale, of New York, and several others.



crystallising at  $32^{\circ}$ , and boiling at  $186^{\circ}$ . It burns with a rich flame, and produces a large quantity of smoke. It is composed, as its name indicates,\* of one equivalent of hydrogen, (1,  $\square$ ) and two of carbon, (12,  $\square\square$ ). Its equivalent by weight is accordingly 13, and its elements are condensed into a space not larger than 1-9th of the volume which they naturally occupy. Water and carbonic acid are the products of its combustion, a third of a measure, (one equivalent by measure,) requiring two and a half measures of oxygen, (five equivalents,) one of which combines with the hydrogen and forms the water, while the rest goes to the carbon and converts it into carbonic acid.

*Naphthaline* is another compound of carbon and hydrogen, which is obtained, when the distillation of naphtha from coal tar is conducted at a low temperature, condensing in crystals, which are found in the neck of the retort. It may be regarded as a compound of one equivalent of olefiant gas and one of the bi-hydrogenated carbon, as every twenty parts by weight consist of eighteen of carbon and two of hydrogen, according to the analysis of Dr Thomson. It is white when pure, melts at  $180^{\circ}$ , and boils at  $110^{\circ}$ . It is insoluble in cold water, but is dissolved by naphtha, alcohol, and ether, and, in small quantities, by hot water.

Naphthaline combines with sulphuric acid, and forms a compound which has also acid properties; it is termed the *sulphonaphthalic acid*.

#### SECTION V.—OF ALCOHOL.

It will be recollected, that alcohol, ethers, &c., are usually noticed at or near the close of chemical treatises. But as these substances, in common with the vegetable acids, are composed, almost without exception, of carbon, oxygen, and hydrogen, there is a propriety, in point of arrangement, as well as profit to the student, in introducing them in this part of the work.

The equivalent of alcohol, by weight, is 23. It is composed of oxygen 8, (one equivalent,) carbon 12, (two equivalents,) and hydrogen 3, (three equivalents); or of water 9, (one equivalent,) and olefiant gas 14, (two equivalents). By volume, its equivalent is one whole measure. Its specific gravity, when pure, is 0.796. The vapor of alcohol has a specific gravity of 1.597. One hundred cubic inches weigh 48.714 grains.

\* If but one equivalent of each were present, then the name would be *carburetted* hydrogen, for the reason assigned under the head of carburetted hydrogen or olefiant gas. But as two equivalents of carbon are joined to one of hydrogen, the prefix *bi* is highly proper.



Alcohol is a term of alchemical origin, signifying the pure substance of bodies. It is now applied to the pure spirit formed by vinous fermentation, and separated by distillation. Those are mistaken, altogether, who suppose that distillation *forms* the alcohol, as well as separates it. Unless the substances which are destined to yield alcohol, undergo the vinous fermentation, or that intestine, decomposing motion, which results in the formation of vinous spirit, the mere operation of distilling will never afford it. Hence, it follows, that a liquid may possess the power of intoxicating, that has never been subjected to the still. Our domestic wines furnish sufficient proof of the truth of this assertion.

The term fermentation, is generally used to express those changes which dead animal and vegetable matter, in contact with water, undergoes at natural temperatures. When a solution of sugar passes into this state, an intoxicating liquid is produced, such as beer or wine, its properties being modified more or less, by the various mucilaginous matter and other substances which the liquid may contain. This is termed the *vinous* fermentation, as the product, in all cases of this kind, is a portion of alcohol, which imparts to these liquids their characteristic properties, and similar to that which is obtained from the juice of the grape.\* When the liquid becomes sour, the *acetic* fermentation is said to take place, as a quantity of acetic acid or vinegar is then formed. All moist vegetable and animal substances are subject to a still farther change, from the spontaneous reaction of their elements, which is termed the *putrefactive* fermentation, when a large quantity of offensive gases is always disengaged.

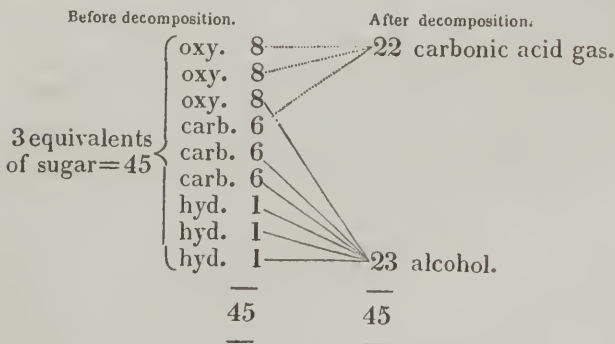
To see the changes that take place during the vinous fermentation, put a solution of three or four ounces of sugar in four times its weight of water, into a glass retort or flask; add a small quantity of yeast, which may be procured from any brewery, and put the beak of the retort, (or the extremity of a bent tube fitted to the flask, if it is preferred,) under the mercury in a pneumatic trough, a jar being placed above it to collect the gas that is disengaged. If there is no proper apparatus for fixing the mercurial jar, the experiment may be performed with a water trough; but in this case, a considerable portion of the gas, which is disengaged, will be ab-

Exp.

\* If sugar, or the grains, be the source of alcohol, there is a necessity, imposed by the very constitution of things, that the characteristic features of these substances be destroyed. Hence, fermentation may be fitly denominated a *new creation*.

sorbed. The solution should be placed where it will be exposed to a temperature between  $70^{\circ}$  and  $80^{\circ}$ ; it will soon become turbid, and a quantity of carbonic acid gas be disengaged, (which may be recognised by the tests already mentioned); a frothy scum will collect on its surface, and, in a few days, the liquid will be clear. On examining it, in this state, the sugar will be found to have been in a great measure decomposed, and, also, part of the yeast; and the alcohol which it now contains, with the carbonic acid disengaged, will equal, in weight, the sugar which has disappeared. It is by the decomposition of the sugar, then, that these new products are formed, the yeast causing the intestine changes, by which they are produced, to commence, though the precise mode in which it effects this has not been explained.

The manner in which the alcohol and carbonic acid are produced by the decomposition of the sugar, may be easily illustrated. Every 15 parts of sugar (one equivalent) are composed of 8 parts of oxygen, 6 of carbon, and one of hydrogen, or of one equivalent of each of these elements; and three equivalents of this substance give one of alcohol and one of carbonic acid. The elementary composition of three equivalents of sugar, and the manner in which their component parts arrange themselves, will be readily understood by the following diagram.



One equivalent of alcohol is accordingly 23, and it consists of one equivalent of oxygen, two of carbon, and three of hydrogen. It is usually represented as a compound of one measure of watery vapor, (one equivalent,) and an equal bulk of olefiant gas, (two equivalents,) condensed into a volume no larger than one measure, which is therefore its equivalent, by volume. The following diagram shows the manner in which its elements must be arranged according to this opinion.

23 □ alcohol.	{ oxy. 8	.....	9 □ water.
	{ carb. 6	.....	
	{ carb. 6	.....	
	{ hyd. 1	.....	
	{ hyd. 1	.....	7 □ olefiant gas.
	{ hyd. 1	.....	7 □ olefiant gas.
	—		—
	23		23
	—		—

*Yeast* or *ferment* is the term applied to the substance which separates in small, viscid flocculi, from all solutions undergoing the vinous fermentation, and capable of commencing the same fermentation in liquids similar to those by which it is produced. When dried by a gentle heat, it may be preserved without change, but if heated in boiling water, it is no longer capable of exciting the vinous fermentation. It consists of carbon, oxygen, hydrogen, and nitrogen.

If the fermented liquor is exposed to heat in a retort, the alcohol, from its greater volatility, rises first, carrying a portion of water along with it, and may be condensed in a receiver. By stopping the distillation when the last portions of alcohol have been disengaged, it may be separated, not only from the mucilaginous matters which it may have been previously combined with, but also from the greater part of the water. It is in this manner that brandy is prepared from wine, and whiskey from the fermented infusion of malt. The process may be imitated, on the small scale, by heating half a bottle of wine or ale in a glass retort, and continuing the distillation as long as any inflammable liquid comes over. When ale

Exp.

or porter is employed, a capacious retort must be used, as the large quantity of carbonic acid which they contain, is disengaged by the heat, and would cause part of the fluid to run over into the receiver, if the distillation were conducted in a small retort.

By repeating the distillation, the alcohol may be separated from an additional quantity of water which is left in the retort, but it is impossible in this manner to separate the whole of the water, however often the distillation may be repeated.

To prepare *absolute* alcohol, (as it is termed, when completely deprived of water,) heat subcarbonate of potash to the temperature of 300°, and add it to spirit of wine, in a glass bottle. Shake the mixture well, and then allow it to remain at rest for some time. If a sufficient quantity of the alkaline carbonate has been added, the liquid will divide into two parts; that which floats above, is the alcohol deprived of most of the water previously combined with it, while that below,

Expt

consists of the water which has dissolved the subcarbonate, and formed a dense, oily-looking fluid. The alcohol is decanted, or drawn off with a syphon, and more of the hot subcarbonate is added, till it is no longer moistened. On decanting it again, and distilling it with a gentle heat, the alcohol is obtained free from water. Nearly half a pound of subcarbonate is required for every pint of rectified spirit of wine.

Other substances are occasionally employed to separate the water from the alcohol, instead of the subcarbonate of potash, as the chloride of calcium, lime, barytes, and alumina. Very strong alcohol may also be procured, by suspending a bladder full of spirit of wine in the air, the water passing through the coats of the bladder and evaporating from its surface, while the alcohol is retained.\*

Mr Graham's process for preparing absolute alcohol, may be easily conducted by those who have an airpump. A shallow glass or earthen bason is filled with quick lime in coarse powder, and another, nearly full of spirit of wine, put over  
 Exp. it; they are then placed on the plate of the airpump, and the air exhausted till the liquid appears as if it were beginning to boil. Both the water and alcohol evaporate at first, the watery vapor being absorbed by the lime, which does not affect the vapor of the alcohol; but water does not remain in combination with alcohol, unless covered by an atmosphere of its own vapor, and as this is condensed by the lime as speedily as it is formed, the water continues to evaporate till it is completely removed, which usually requires three or four days, while the alcohol is prevented from evaporating, by the pressure of its own vapor.

The quantity of alcohol in spirituous liquids may be ascertained, by adding hot subcarbonate of potash to them, in the manner already mentioned, after precipitating the coloring and mucilaginous matters, by dropping into them a strong solution of the acetate of lead, or agitating them with the protoxyde of lead in fine powder. By operating in this manner, Mr Brande proved, satisfactorily, that alcohol really exists in fermented liquors, and is not formed by any reaction taking place between their elements when heat is applied.† A table of the results of Mr Brande's experiments, and the quantity of real

\* It has been observed in the laboratory of the Royal Institution, that diluted spirit becomes stronger by being kept in vessels closed by bladders; showing that the alcoholic vapor passes through these membranous substances, less freely than aqueous vapor.

† This is a complete refutation of the notion, that fermented liquors do not contain alcohol, and that distillation is essential to its formation. This fallacy has been advocated from the press.

alcohol in a variety of wines and other fermented liquors, is published in his Manual of Chemistry.

Alcohol and water combine in every proportion, condensation and an evolution of heat usually accompanying the combination, which may easily be seen by filling a tube ten or twelve inches long and already half full of water with alcohol, and shaking them together. Thenard found, that when alcohol of the specific gravity of 0.9707 is mixed with an equal quantity of pure water, instead of any condensation taking place, they occupy a larger volume after they have combined than before. With an equal bulk of water, it forms what is usually termed *proof spirit*, the specific gravity of which is 0.917. The *proof spirit* used for pharmaceutical purposes has usually a specific gravity of 0.930, and contains 43 per cent., of real alcohol. The following table, by Lowitz, shows the quantity of real alcohol in diluted alcohol, at different densities, and will be found useful in making experiments with this liquid. Though I have adopted the number 0.796 to represent the specific gravity of absolute alcohol, perhaps Lowitz is the only one who has obtained it so light. According to the London college its specific gravity is 0.815. It is seldom prepared for commercial purposes, with a less specific gravity than 0.835 or 0.830.

*Lowitz's Table, showing the quantity of absolute alcohol in 100 parts of spirit at different densities.*

Specific gravity at 60.	Alcohol in 100pts	Specific gravity at 60.	Alcohol in 100pts	Specific gravity at 60.	Alcohol in 100pts	Specific gravity at 60.	Alcohol in 100pts	Specific gravity at 60.	Alcohol in 100pts
0.796	100	0.848	80	0.896	60	0.939	40	0.974	20
0.798	99	0.851	79	0.898	59	0.941	39	0.975	19
0.801	98	0.853	78	0.900	58	0.943	38	0.977	18
0.804	97	0.855	77	0.902	57	0.945	37	0.978	17
0.807	96	0.857	76	0.904	56	0.947	36	0.979	16
0.809	95	0.860	75	0.906	55	0.949	35	0.981	15
0.812	94	0.863	74	0.908	54	0.951	34	0.982	14
0.815	93	0.865	73	0.910	53	0.953	33	0.984	13
0.817	92	0.867	72	0.912	52	0.955	32	0.986	12
0.820	91	0.870	71	0.915	51	0.957	31	0.987	11
0.822	90	0.872	70	0.917	50	0.958	30	0.988	10
0.825	89	0.874	69	0.920	49	0.960	29	0.989	9
0.827	88	0.878	68	0.922	48	0.962	28	0.990	8
0.830	87	0.879	67	0.924	47	0.963	27	0.991	7
0.832	86	0.881	66	0.926	46	0.965	26	0.992	6
0.835	85	0.883	65	0.928	45	0.967	25	0.994	5
0.838	84	0.886	64	0.930	44	0.968	24	0.995	4
0.840	83	0.889	63	0.933	43	0.970	23	0.997	3
0.843	82	0.891	62	0.935	42	0.972	22	0.998	2
0.846	81	0.893	61	0.937	41	0.973	21	0.999	1



Absolute alcohol is a transparent and colorless liquid, having a fragrant odor, and a hot, pungent taste. It is not only lighter, but also much more volatile than water; when its specific gravity is 0.810, it boils at  $175^{\circ}$ . It is much more expansible than water by heat, and has been frozen by the cold produced by the evaporation of liquid sulphurous acid, in vacuo. Mr Walker exposed it to a temperature of  $91^{\circ}$ , without freezing it.

Alcohol burns with a pale-blue, lambent flame, which becomes yellow on diluting it with water, and every 23 parts, in burning, consume 48 of oxygen. The products of the combustion are water and carbonic acid. A number of salts, when mixed or dissolved in it, impart particular colors to the flame, especially the muriates of strontites, barytes, and copper; the first of which gives a deep red, the second a yellow with green tinge, and the last a fine, bright green. These varieties, in shade, afford some evidence, by which to distinguish between those salts. With a great number of substances, alcohol combines in definite proportions, according to Mr Graham, forming a peculiar class of compounds, which he has termed *alcoates*. It is decomposed by potassium and sodium, which attract oxygen from it, and disengage hydrogen gas.

Alcohol is an agent that is constantly employed in the laboratory, for affording a steady and powerful heat during its combustion,\* and in a great number of operations, where it is used as a solvent, or to afford particular compounds by the reaction of its elements on other substances. It is particularly useful in dissolving resins, camphor, volatile oils, vegetable acids, and alkalis, and many salts and other substances which are insoluble in water, enabling us to separate them from other bodies with which they may be mixed or combined. From the large quantity of hydrogen and carbon which it contains, it has been occasionally employed to deoxidate particular substances, and from the rapidity with which it evaporates, it is sometimes employed to produce cold.

Notwithstanding all these advantages, it is more than probable that the world has gained nothing, really, by the introduction of alcohol; and it would not be difficult to show, that the best interests of society would be promoted, if all nations pro-

\* Professor Hare recommends the addition of a seventh part of spirits of turpentine, to the alcohol intended for spirit lamps. He affirms that the heat is much increased, without any disadvantageous circumstance, attendant on the admixture. Mr Guthrie, a contributor to Silliman's Journal, has made an article which he calls *pure spirits or oil of turpentine*, which he burns in mixture with alcohol, affirming that the combustion does not leave those resinous spots on the wick, nor cause the scintillations that are noticed, when the common spirit of turpentine is used.

hibited its formation, by the most rigorous statutes. But if these views be too extravagant in the estimation of the many, I hope to see the day, at least, when alcohol will be restricted to the laboratory and to the useful arts. It is my design, ere long, to prepare a *Manual of Pharmacy*, in which the nature, compounds, and uses of medicinal substances will be examined, apart from the pernicious influence of alcohol.

In the mean time, it is proper to say that alcohol is used in medical practice, in various states of dilution; and if a man should be so straitened in remedial aids, that nothing else could be had, then, and then only, ought he to employ this substance. There is not, however, so much ground for objection to its external, as to its internal use, especially when combined with matters that unfit it for the latter mode of application.

In relation to the poisonous effects of alcohol, as seen in drunkenness, it is necessary to say, that chemistry furnishes no antidote, and that the treatment must be purely medical.

#### SECTION VI.—OF ETHERS.

In close relation to alcohol, we very properly place the different articles that have been denominated *ethers*. By the term ether, we generally mean a volatile fluid, produced by the action of an acid on alcohol; and the resulting compound must necessarily vary with the acid employed. The most important ethers are the sulphuric and nitric. I perceive, by a recent publication in Professor Silliman's Journal, that *chloric* ether is likely to acquire considerable importance, as a medicinal agent.

##### I. Of Sulphuric Ether.

Basil Valentine was the first person who noticed the production of ether from alcohol; and, in his day, this term was applied, exclusively, to the volatile fluid, formed by the action of sulphuric acid on alcohol.

Experiments, conducted with much care, have shown that sulphuric or vitriolic ether, is a compound of 8 oxygen, (one equivalent,) 24 carbon, (four equivalents,) and 5 hydrogen, (five equivalents); or, water 9, and olefiant gas 28; making its equivalent, by weight, 37. Its equivalent, by volume, is one whole measure. When perfectly pure, its specific gravity, according to Lowitz, is 0.632, but we seldom get it below 0.720, and not often of that strength. It boils at  $98^{\circ}$ , in the open air, and in vacuo at  $40^{\circ}$ . One hundred cubic inches weigh 78.36 grains.

To prepare sulphuric ether, equal weights of sulphuric acid and alcohol are exposed to heat in a plain glass retort, pouring in the alcohol first, and then the acid, by a bent glass funnel, and adjusting the retort in a sand bath already heated to the temperature of  $200^{\circ}$ . The acid and the alcohol should be well mixed, previously, by shaking them together in the retort, when the temperature rises considerably; the receiver should be rather larger in proportion to the size of the retort, and tubulated, to convey away the atmospheric air, and any other gaseous products that may be formed towards the close of the operation. The neck of the receiver should fit closely to the neck of the retort, and the joint be rendered tight, by tying it round with a piece of linen or cotton cloth, spread over with a paste made of flour or linseed meal. When a large quantity of materials is operated on, a bent tube should be made to pass into a second receiver, which should be kept cold by placing it in a jar or basin full of water or ice; the tube must not fit tightly to the neck of the receiver, but allow any gas to be freely disengaged. The first receiver should be covered with a piece of linen or cotton cloth, that it may be more easily kept cold; ice or snow should always be used for that end, when they can be procured.

The distillation is generally continued, till a quantity of liquid has come over, equal to one half of the alcohol employed. More ether is said to be obtained, when it is kept constantly boiling, than at a lower temperature, though this has been disputed; the retort should not be filled more than half full, and great attention must be paid to the heat applied during the whole of the operation, as the mixture is apt to boil over when urged with too strong a fire. By adding half the quantity of alcohol employed at first, an additional quantity of ether may be obtained, and this may be repeated again; the acid soon becomes so diluted, however, that it is unable to produce any more.\*







The ether that condenses in the receiver, is never obtained pure at first, being always mixed with a little alcohol that distils over, unaltered, and some sulphurous acid produced by the decomposition of part of the sulphuric acid. To remove these, it is mixed with fused potash, taking five or six grains for every ounce of alcohol employed, and distilled again from a retort, with a very gentle heat, till five or six-sevenths of it shall have passed over. The potash retains the sulphurous acid along with

\* Mr C. Wesener, a practical chemist, living near Philadelphia, makes ether by the application of steam. The process is more safe and economical, than that in general use.

some water and alcohol; and to separate the alcohol completely, the product of the second distillation may be shaken with about three-fourths of its bulk of water, which combines with all the alcohol and a little ether, the most of the ether separating from it in a short time and floating above. It is then decanted, and kept in bottles with ground glass stopples.

On the small scale, an ounce or two of alcohol with as much sulphuric acid, by weight, will be sufficient to show the process, condensing the product in a Florence flask.

The theory of the action has already been adverted to, being precisely similar to what takes place in the first stage for the preparation of olefiant gas. (See the diagrams representing the composition of alcohol.) In the present instance, however, a much smaller quantity of sulphuric acid being used, a larger quantity of ether is formed. The sulphuric acid, by its powerful affinity for moisture, takes up a large portion of the water present, in one equivalent of alcohol, the rest being dissipated by the heat. This equivalent of alcohol, thus deprived of water, is reduced to two equivalents of olefiant gas, 14, which combine with one undecomposed equivalent of alcohol, (consisting of 14 olefiant gas, and 9 water,) and give one equivalent of ether, 37. Hence, there will be a large quantity obtained; and it is obvious, also, that there must be twice as much olefiant gas combined with a given quantity of water in ether, as there is in alcohol.

	Water.		Olefiant gas.	
Equivalents by weight	9	+	14	= 23 alcohol.
Equivalents by weight	9	+	28	= 37 ether.
Equivalents by volume		+		=  alcohol.
Equivalents by volume		+		=  ether.

And by examining the corresponding equivalents, by weight and by volume, we see that the vapor of ether must be heavier than that of alcohol, containing twice as much olefiant gas as exists in alcohol, condensed within the same space.

If the distillation is continued, after a quantity of ether equal to one half of the alcohol employed, has distilled over, a peculiar substance begins to rise in the retort, and condense in the form of an oily liquid, which has been called the *oil of wine*. It is composed of one equivalent of sulphuric acid and four of olefiant gas, and has no acid properties, the sulphuric acid which it contains being completely neutralised by the olefiant gas, a property which we certainly could not suppose it would possess. When exposed to heat, half of the olefiant gas is disengaged, and the compound which remains, is what has been termed *sulphovinic acid*, consisting, accordingly, of one equivalent of sulphuric acid, and two of olefiant gas.



Water can dissolve only a small quantity of ether, but alcohol and ether combine in every proportion. It is very inflammable, and burns with a richer and much more copious flame than alcohol; the products of its combustion are water and carbonic acid.\* A few drops put into a detonating bottle, full of oxygen gas, and immediately corked, speedily diffuse themselves through the gas, and form an inflammable mixture, that  
 Exp. detonates violently on bringing a lighted match to the mouth of the bottle. This is an experiment that should be performed with a very small and strong bottle, as detonating bottles that have not been injured by any other explosive mixtures, are frequently broken by this. When transmitted through a red hot tube, it is decomposed, and gives the same products as alcohol.

From the rapidity with which ether evaporates at natural temperatures, it is often used to produce an intense degree of cold. If a small quantity is poured into a jar, which is immediately covered with a tray, it will speedily evaporate,  
 Exp. and on applying a lighted candle to the mouth of the jar, it will be found to be full of inflammable vapor. If a larger quantity of ether is put into an open jar, and a coil of thin platina wire, heated to redness in a spirit lamp, is suspended over it at a particular distance, (which is easily found on trying the experiment,) instead of becoming cold, it remains red hot, till the whole of the ether is consumed. The platina, being a bad conductor of caloric, does not part with its excess of heat, so rapidly as any other metal would do, and coming in contact with the vapor of ether and air, in the jar, the high temperature of the metal causes a slow combustion to take place, by which a sufficient degree of heat is produced, to maintain it in a state of incandescence. The evaporation of ether, and the consequent production of cold, (or rather the abstraction of heat,) is accelerated by the removal of atmospheric pressure. Thus, if we place a thermometer tube containing some water, in a thin glass vessel, partly filled with ether, and fix the whole under the receiver of an airpump, a few strokes of the pump will so far exhaust the receiver, as to remove atmospheric pressure sufficiently to allow the evaporation of the ether. In a few minutes, the water in the tube will be thus frozen.

Sulphuric ether is not capable of dissolving so many substances as alcohol; still, however, it is often found useful in

\* In the conversion of one equivalent of ether into water and carbonic acid, by the process of combustion, 12 equivalents of oxygen are consumed.



separating or extracting principles that are insoluble in alcohol or water, more especially in vegetable chemistry. It combines with ammonia, camphor, resins, volatile oils, sulphur, phosphorus, and chloride of gold, but has little or no action on the fixed alkalis, earths, common metallic oxydes, and the greater number of the salts.

On account of its tendency to evaporate, at ordinary temperatures, ether should be kept in well stoppered bottles, and the mouths should be covered with bladders. It cannot be prepared on high, mountainous situations, without loss; for, in proportion as the pressure of the atmosphere is removed, just in that proportion does its tendency to fly off, in vapor, augment.

In pouring it from one vessel to another, care should be taken to keep candles and flame of any sort, at a distance, especially in warm weather. I knew a manufacturing chemist, who was burnt very badly by the unexpected firing of a large quantity of ethereal vapor. He had put his carboys of ether in a cool cellar, to have them kept in perfect security; but, by some means, one of them had been broken, and the contents were diffused throughout the cellar. He took a candle in his hand, and entered the cellar, but in an instant he was surrounded by an atmosphere of flame. The consequences of this accident are so durable, that he will carry them to his grave.

The evaporating property of ether is made use of, to remove drops of water that obstinately adhere to the inside of vials. A little ether is poured in, the vial shaken, and then exposed to a current of air. The ether, in evaporating, carries off the water, and the bottle is completely dried in a few minutes.

Sulphuric ether is sometimes adulterated, and no test is better than its specific gravity. When free sulphuric acid is present, the addition of a solution of barytes, will give a white precipitate, and litmus paper will be reddened. Pure ether forms a limpid solution with phosphorus; but if alcohol be present, the solution will be milky.

If ether be kept undisturbed for a long time, a spontaneous decomposition will occur, and acetic acid be formed.

Perfectly pure ether is seldom used in medical practice, and it should also be recollected, that sulphuric ether, and rectified ether, although synonymous in chemical language, are somewhat different. The former being a sulphurous mixture of ether, alcohol, and water; while the latter is the pure article, free of all sorts of admixture.

The preparations of sulphuric ether, employed in medicine, are as follows: the *spirit of sulphuric ether*, which is the pure ether, diluted with twice its weight of rectified spirit; the *com-*

*pound spirit of sulphuric ether*, which differs from the last article only in the addition of a little ethereal oil, and is used as a substitute for *Hoffman's anodyne liquor*, the precise composition of which is not well understood; next, we have the *aromatic spirit of ether*, which differs from the sulphuric ether only in the addition of several spices.

The spirit of sulphuric ether, and pure ether itself, are sometimes employed as external remedies, on account of the rapidity with which they carry off the animal heat, by evaporation. All the preparations of ether are used internally, but, judging from my own experience, after having been engaged in practice for twenty years, they may be dispensed with altogether. During the whole of that period, although employed in hospital as well as private practice, I am very sure, that I have not prescribed a pint of the preparations of sulphuric ether.

Some years ago, a practice obtained among the lads of Philadelphia, of inhaling the vapor of sulphuric ether, by way of sport. A small quantity, placed in a bladder, was almost instantly converted into vapor, by the application of hot water. By means of a tube and stop-cock, the gas could be easily inhaled. In some instances, the experiment excited mere playfulness and sprightly movements; but in several cases, delirium, and even phrenitis was induced, which terminated fatally.

The vaporisation of ether by heat, may be shown in a moment. Let a bladder, into which a teaspoonful of good ether has been placed, be made perfectly tight, by means of a  
 Exp. string fastened round its neck. Then lay the bladder on some hot water, and presently it will be distended to its full size. At this moment, remove the bladder and cause a stream of cold water to fall on it. In an instant, there will be a complete collapse, owing to the sudden condensation of the ethereal vapor.

## 2. Of Nitric Ether.

The best method of preparing nitric ether, is by mixing equal weights of alcohol and the strong, fuming nitric acid, prepared by distillation from two parts, by weight, of sulphuric acid with three of nitre. The large quantity of nitrous acid which it contains, reacts on the alcohol, and, in a day or two, converts it into ether, which floats on the top of the remaining liquid, and may easily be removed by a small syphon. Pure nitrous acid, prepared by distillation from the nitrate of lead,

would do still better, but it is not so easily obtained. Two or three ounces of alcohol will be sufficient to show the process; the alcohol is put into a bottle first, and small quantities of the acid gradually poured into it through a funnel with a long stem, which passes to the bottom of the bottle, which is to be well shaken after each addition of acid, and then placed in cold water, to prevent violent reaction. A drachm or two of the acid may be added, every quarter of an hour, in this manner, till it is all mixed with the alcohol. The bottle should be provided with a conical stopple, to allow the gas that accumulates in it to be disengaged; it is forced up in the same manner as the stopple in Nooth's apparatus, already described, and returns to its place when the excess of gas has passed by it.

Other methods for the preparation of nitric ether have been proposed. The Dublin college directs the alcohol to be mixed with sulphuric acid in a flask, and the mixture to be poured over bruised nitre in a retort, mixing a pound of the acid with nineteen ounces of alcohol, by measure, and using three parts, by weight, of nitre, for every two of sulphuric acid employed. The retort must be placed in a bason of cold water, to prevent the action becoming too violent, and it should not be more than a third full of nitre. It is seldom necessary to apply any heat to commence the action, as is usually recommended. The sulphuric acid combines with the potash, and the nitric and nitrous acids, acting on the alcohol in its nascent state, produce the ether, which must be condensed in a large tubulated receiver, kept very cold; when a large quantity is required, a second tubulated receiver should be connected with it, and the gaseous products allowed to escape by another bent tube. The first method of preparing nitric ether will be found preferable, however, as the reaction often becomes very turbulent, when this process is adopted, though every precaution is taken to prevent it.

In all experiments with nitric acid and alcohol, great care must be taken, not to mix a large quantity of acid with the alcohol, at once, as the gaseous products that are immediately produced, are apt to throw out the whole of the mixture, with explosive violence. Though a small quantity of acid may be added to a large quantity of alcohol, without any particular action being observed, a small quantity of alcohol cannot be mixed with a large quantity of acid, without being completely decomposed, as the particles of the alcohol are surrounded by the acid on every side, which affords oxygen more readily to the inflammable elements that enter into its composition. To

see the truth of this, all that is necessary is to pour a little alcohol and acid into different glasses, and pour a few drops  
 Exp. of the acid into the alcohol, and then of the alcohol into the acid, when the appearances described, will be observed.

Nitric ether always contains a little acid, as it is procured at first, which may be removed by potash or lime. It has a very pale lemon-yellow color; a pleasant smell, similar to that of apples; and a strong, penetrating taste. It is heavier and more volatile than sulphuric ether, burns with a rich flame, and soon becomes acid by keeping. When it is purified by distillation, the operation should always be carried on with a very gentle heat, as it is decomposed when distilled quickly at a higher temperature.

Its atomic composition is still disputed, but it is admitted to contain a portion of nitrogen, in addition to the other elements which exist in sulphuric ether.

### 3. *Spirit of Nitric Ether, or Sweet Spirit of Nitre.*

The spirit of nitric ether is a compound of nitric ether and alcohol, which is prepared by mixing nitric or nitrous acids with a larger quantity of alcohol than is used in the process for ether, and distilling the mixture in a glass retort. The Edinburgh college directs one pound of their nitrous acid to be mixed with three of alcohol, and distilled with a heat not exceeding  $180^{\circ}$ , till a quantity of liquid has been obtained, equal to the alcohol employed. Dr Duncan says, the distillation may be commenced whenever the materials have been mixed; the precautions already pointed out must be attended to, and the receiver kept cold in the usual manner.

Mr Brande, in his Manual of Pharmacy, directs three ounces, by weight, of nitric acid to be added gradually to two pints of rectified spirit. The temperature must not be permitted to rise above  $120^{\circ}$ , and this end may be gained, by immersing the vessel in cold water. After the mixture is perfect, a gentle heat is to be applied, and twenty-four fluid ounces distilled off.

The United States' Pharmacopœia gives the following direction. Take of nitrate of potash, in powder, two pounds; sulphuric acid, one pound and a half; alcohol, nine pints and a half; diluted alcohol, a pint; carbonate of potash, an ounce. Mix the nitrate of potash and the acid in a glass retort, and, having gradually poured in the alcohol, digest with a gentle heat for two hours; then raise the heat, and distil a gallon. To the distilled liquor, add the diluted alcohol and carbonate of potash, and again distil a gallon.



The *spirit of nitric ether*, or *sweet spirit of nitre*, has been analysed by Thenard, and appears to consist of five equivalents of carbon, 30, six of oxygen, 48, eight of hydrogen, 8, and one of nitrogen, 14, in 100 parts.

This article has a peculiar, but grateful odor, a pungent and rather sweet taste. If kept for several months, it acquires a slight degree of acidity. It readily dissolves in water and alcohol. If tincture of guaiacum be dropped into it, a fine and peculiar blue tint appears, passing into various shades of green, while the mixture continues to be transparent. Exp. If water be added, a blue or green precipitate falls, which at length becomes brown.

The sweet spirit of nitre is used, internally, in the practice of medicine.

Agreeably to our definition of an ether, at the opening of this section, it will be seen, that there are almost as many ethers as we have acids. Thus, there are the muriatic, hydriodic, acetic, phosphoric, and arsenic ethers, with others which need not be enumerated. They are formed on the same principles with the ethers already considered, and, therefore, shall not further occupy our attention, in this place.

#### 4. Of *Pyroxilic and Pyroacetic Spirit*.

Both these compounds consist of carbon, oxygen, and hydrogen; they are lighter than water, volatile and inflammable. The *pyroxilic spirit* is formed during the preparation of acetic acid, by exposing wood to heat in close vessels. It burns with a blue flame, leaving no residuum, and its specific gravity is 0.828. *Pyroacetic spirit*, or *ether*, as it has been termed, is formed when the metallic acetates are exposed to heat, in the preparation of strong acetic acid. It is lighter than the *pyroxilic spirit*, and is easily distinguished from it, by burning with a dense flame, and combining in all proportions with oil of turpentine. Neither of these substances can be prepared, easily, on the small scale; but *pyroxilic spirit* is frequently employed as a substitute for alcohol in the spirit lamp, when it can be procured, and affords a very steady heat during its combustion, without producing any smoke.

#### SECTION VII.—OF THE BI-CARBURET OF NITROGEN, OR CYANOGEN GAS.

I place the names of this compound, as I find them in Dr Turner's book, and not in the reversed form in which Dr Reid



has them. The *bi-carburet of nitrogen* is the true chemical name, in conformity with the spirit of the modern nomenclature; as it is composed, by weight, of two equivalents of carbon, 12, and one of nitrogen, 14, making its combining number 26. The two equivalents of carbon require the use of the prefix *bi*, as in other similar cases.

As the word *cyanogen* is more convenient for common use, than the *bi-carburet of nitrogen*, and as it was selected by Guy Lussac, (who discovered it in 1815,) on account of its derivative meaning, viz. *a producer of blue*, I shall employ it in the course of these remarks.\*

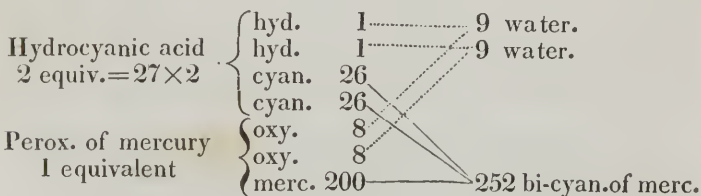
The equivalent of cyanogen, by volume, is one whole measure, and its specific gravity 1.804; one hundred cubic inches weigh 55.06 grains. According to Mr Faraday, it may be liquefied by a pressure of about four atmospheres. It is obtained in the gaseous form, by exposing the *bi-cyanide of mercury*, (a compound of metallic mercury and cyanogen,) to heat in a small glass retort, and collecting the gas that is disengaged, over the mercurial trough. Green glass retorts should be taken when they can be obtained, or an iron tube retort; an ounce of the *bi-cyanide* gives a considerable quantity of gas, and when only a very small quantity is required, it may be procured by heating a little of the *bi-cyanide* in a bent tube with a spirit lamp.

In conducting this process, the *bi-cyanide* employed should not fill the retort more than half full, and be perfectly dry. The heat applied should be sufficient to expel the cyanogen slowly and steadily, as it is liable to be decomposed by a higher temperature. The mercury is volatilised as the gas is disengaged, the *bi-cyanide* melting and becoming black before it is decomposed. Towards the end of the process, a small quantity of nitrogen gas is disengaged, arising from the decomposition of a part of the cyanogen, the carbon, previously in combination with it, remaining in the form of a light porous mass, very like lampblack.

The *bi-cyanide of mercury* is prepared, by boiling pure Prussian blue with one and a half times its weight of the peroxyde of mercury, in an earthen evaporating bason, reducing both to a fine powder, previously, and mixing them with fifteen or twenty parts of water. The Prussian blue is obtained sufficiently pure for this purpose, by digesting the Prussian blue of commerce in muriatic acid, diluted with ten times its bulk of water, which removes the alumina and other foreign matters it

\* It has been also called *Prussine gas*, from Prussian blue, in which it exists as a component part.

usually contains; after this, it must be repeatedly washed with water, till the excess of acid has been removed. The precise nature of Prussian blue has not yet been satisfactorily demonstrated, but it contains hydrocyanic acid, (a compound of cyanogen and hydrogen,) and peroxyde of iron, whatever other ingredients may be present; and in this process, the acid combines with the peroxyde of mercury, forming hydrocyanate of mercury, two equivalents of it uniting with one of the oxyde, while the peroxyde of iron is set at liberty. The hydrocyanate, remaining in solution, is separated by filtration, and on evaporating the liquid, till a pellicle appears on its surface, crystals of the bicyanide of mercury are obtained when it cools; the two equivalents of hydrogen in the two of hydrocyanic acid, unite with the oxygen contained in the peroxyde, while the cyanogen goes to the metallic mercury, as may be seen by the following diagram.



The method of preparing Prussian blue, will be described under ferrocyanate of iron.

Rather more peroxyde of mercury is used, than is necessary to combine with the hydrocyanic acid, that the iron may be completely separated; and, as any excess is dissolved, even when in a considerably greater proportion, it is better, in general, to add some pure hydrocyanic or muriatic acid till the solution is neutralised, small scaly looking crystals being formed, instead of long slender prisms, when this is not attended to, which do not give any cyanogen, on exposing them to heat. They are composed of the bi-cyanide and peroxyde of mercury, and when decomposed by heat, the oxygen of the oxyde, uniting with the carbon of the cyanogen, disengages nitrogen gas, and forms carbonic acid and carbonic oxyde. Little or no cyanogen is obtained even from the neutral cyanide, when it is moist, the hydrogen of the water converting one part of the cyanogen into prussic acid, and forming ammonia with the nitrogen of another portion, the oxygen combining at the same time with the carbon and producing carbonic acid.

Cyanogen gas is transparent and colorless, and has a peculiar, pungent odor. It cannot support combustion, but burns with

a beautiful, purple flame; a small jar, filled with it at the mercurial trough, will be sufficient to show its inflammability, or it may be inflamed as it is disengaged from a small tube, the open extremity having been drawn out to a small point at the blow-pipe. It consumes twice its volume of oxygen during its combustion, each of the two equivalents of carbon which it contains in a condensed state, taking two of oxygen and forming carbonic acid gas, while the nitrogen is disengaged. When it is mixed with pure oxygen in this proportion, the mixture must be put into the detonating bottle before it is inflamed. Cyanogen can bear a much higher temperature than most other compound gases, without decomposition; water absorbs more than four times its volume of this gas, and alcohol dissolves nearly six times as much.

The watery solution reddens litmus paper, but this effect is not produced by the cyanogen, but by the formation of acids, resulting from the decomposition of water. Thus, the hydrogen of the water combines with a portion of the cyanogen, forming hydrocyanic acid, while the oxygen of the water combines with another portion of cyanogen, and gives rise to cyanic acid.

All the compounds of cyanogen that are not acid, are called *cyanurets* or *cyanides*, the latter term being generally preferred in this country. Cyanogen, although a compound body, combines with the simple, non-metallic substances, generally, and with several of the metals. When potassium is heated in cyanogen gas, a very energetic action ensues, the metal becomes incandescent, and a cyanuret of potassium is generated. It is also capable of uniting with the alkalis, but the compounds, thence resulting, are by no means permanent. A strong red heat will effect the decomposition of all the compounds, of which cyanogen makes a part.

We have seen, that when cyanogen is absorbed by water, two acids are formed, viz., the *cyanic* and *hydrocyanic*; and if the water hold in solution an alkali or alkaline earth, these acids will unite with it, and form a cyanate and a hydrocyanate.

These salts are not easily separated from each other, and it is difficult to obtain the cyanic acid from any base with which it is combined, as it is very apt to be changed into carbonic acid, by the reaction of its elements, if water be present.\*

\* The elements of cyanic acid are carbon, nitrogen, and oxygen; the union of the first and last of these, gives rise to carbonic acid, and the hydrogen of some decomposed water joins the nitrogen, and ammonia is the result.

According to M. Liebig, cyanic acid may be obtained in a free state, by transmitting sulphuretted hydrogen gas through water, in which cyanate of silver is suspended; but the operation should be discontinued before all the cyanate is decomposed; otherwise, the free sulphuretted hydrogen would react on the cyanic acid, and effect its decomposition. The acid thus formed, is permanent only a few hours.

*Cyanic acid* is the same with that which was termed, by M. Liebig, *fulminic acid*. The compound called fulminating mercury, (formerly said to consist of fulminic acid and mercury,) was carefully analysed by Gay Lussac and Liebig, who ascertained that the acid in this compound consists of one equivalent, or 26, of cyanogen, and one equivalent, or 8, of oxygen, making a true cyanic acid. This acid has not been applied to any useful purpose. The cyanates of silver and mercury are its most important compounds; they detonate violently, by heat or percussion, and when mixed with some of the acids.

*Hydrocyanic, or prussic acid*, is a compound of hydrogen and cyanogen. It was discovered in the year 1782, by Scheele, who called it prussic acid. Berthollet was the first to ascertain that it was composed of carbon, nitrogen, and hydrogen. Gay Lussac first obtained it in a state of purity, and showed that the compound cyanogen, was acidified by hydrogen, or rather that the union of these substances constituted hydrocyanic acid. As one equivalent, by weight, of cyanogen, 26, is joined to one equivalent of hydrogen, 1, it is plain, that the combining number of this acid must be 27. By volume, its equivalent is two measures, and the weight of 100 cubic inches, 28.50 grains. The specific gravity of the liquid acid is 0.7508, at the ordinary temperature. Its boiling point is  $80^{\circ}$ , and it freezes at zero.

The cyanide of mercury, already spoken of, was formerly called prussiate of mercury, because it was supposed to be a compound of that metal and the *prussic acid* of Scheele. Until the name *hydrocyanic acid* was introduced, chemists were ignorant of the real nature of the acid; hence, we find that prussic acid was one of the mysterious bodies, that refused, for a long time, to yield any satisfactory results to chemical analysis.

Hydrocyanic acid is a very powerful poison, and more speedy in its operation than almost any other. Even the fumes of the acid, when incautiously inhaled, have occasioned stupor, and hence, the operator is obliged to be exceedingly careful, in the preparation of this compound.

There are many processes mentioned in the books, for the formation of this acid, but only two or three will be noticed.

The first, by Vauquelin, is as follows. A glass tube, about 18



inches long, and a half inch wide, is filled with fragments of the bi-cyanide of mercury, and placed in a horizontal direction. A current of sulphuretted hydrogen gas is then passed over these fragments, and in its course, a double decomposition ensues. The hydrogen of the sulphuretted hydrogen seizes the cyanogen of the bi-cyanide, and hydrocyanic acid is the result; while the sulphur of the sulphuretted hydrogen joins the mercury of the bi-cyanide, and a black sulphuret of mercury is produced. The progress of the sulphuretted hydrogen gas may be traced along the tube, by the change of color in the bi-cyanide, and when all of this has become black, the experiment may be closed. We have then to expel the hydrocyanic acid from the tube, by means of a gentle heat, and collect it in a cool receiver. This process is easy, elegant, and productive.

One equivalent of the bi-cyanide, requires two equivalents of sulphuretted hydrogen for its complete decomposition, and two equivalents of hydrocyanic acid and one of the bi-sulphuret of mercury are obtained; the following diagram gives a view of the action that takes place.

Before decomposition.		After decomposition.		
Sulphuretted hydrogen 2 equiv. or $17 \times 2 = 34$	{ hyd. 1	27 hydrocyanic acid.	27 hydrocyanic acid.	
	{ hyd. 1			
	{ sulph. 16	232 bi-sulphuret of merc.	232 bi-sulphuret of merc.	
	{ sulph. 16			
Bi-cyanide of merc. 252	{ cyan. 26			
	{ cyan. 26			
	{ merc. 200			
<hr/> 286 <hr/>	<hr/> 286 <hr/>	<hr/> 286 <hr/>		

The hydrocyanic acid obtained in this manner, is very strong and pure, and equal in weight, when carefully collected, to about a fifth part of the bi-cyanide employed.

The following formula is given in the United States' Pharmacopœia, and is there stated to yield an acid of the same density with that of Scheele. It is an easier process than the one just detailed.

Take of the cyanuret of mercury (the same with the bi-cyanide) one ounce, distilled water eight ounces and a half, carbonate of lead a sufficient quantity. Dissolve the cyanuret, with a gentle heat, in the distilled water; then pass sulphuretted hydrogen gas through the solution, in a proper vessel, till it is fully saturated, and filter through paper. To the filtered liquor, add a quantity of carbonate of lead, more than sufficient to sat-



urate any excess of sulphuretted hydrogen gas, and shake the mixture well. Lastly, filter once more through paper.

The rationale of the first part of this process is the same with that in the method of Vauquelin, but as the hydrocyanic acid is held in solution and a portion of sulphuretted hydrogen gas is present, it is necessary to add the carbonate of lead, in order to remove the gas, which combines with the lead and forms an insoluble sulphuret. The filtered liquor is a weaker acid than that obtained by the process of Vauquelin. Dr Ure's method is precisely the same with that which the United States' Pharmacopœia has adopted.

Brande mentions another mode, which is pursued at the Apothecaries' Hall, in London. One part of the bi-cyanide is mixed with an equal weight of muriatic acid and six parts of water, and then distilled by a gentle heat. The vapors, as they pass over, are to be condensed by the application of cold, and the process continued, until a quantity is condensed, equal to the bulk of the water employed.

The rationale of this process will be seen by the following diagram.

74 or 2 equiv. of mur. acid	{	hyd. 1	-----	27 hydrocyanic acid.
		hyd. 1	-----	27 hydrocyanic acid.
252 or 1 equiv. bi-cyanide of mercury	{	chlor. 36	-----	272 bi-chloride of mercury.
		chlor. 36	-----	
		cyan. 26	-----	
		cyan. 26	-----	
		merc. 200	-----	
326		326		326

As a considerable quantity of water is also employed in this process, it will be obvious, that the acid obtained must be in a diluted state.\*

\* Since the above notices of the preparation of hydrocyanic acid were written, I have met with the following method, as given by M. Tilloy, in the *Journal de Chimie Medicale*. He says, that the acid, thus made, can be kept for several years unchanged, an object which is certainly very desirable.

Take of bi-cyanide of mercury one part, distilled water 4 parts, and alcohol (of 36° Baume) 4 parts. Dissolve the first in water, by the aid of heat, then mix it with the alcohol, and add a very slight excess of hydrosulphuric acid. After this, throw in subcarbonate of lead, and agitate the vessel several times in a water bath, so as to get all the alcohol saturated with the hydrocyanic acid.

Concentrated hydrocyanic acid is speedily decomposed by the reaction of its own elements, and a diluted acid is generally employed for most chemical and medical purposes. In this state, it may be procured much more easily than by any of the preceding processes. The ferrocyanate of potash is dissolved in a small quantity of water; and half its weight of sulphuric acid, previously diluted with twice its bulk of water and allowed to cool, is poured over it in a small glass retort connected with a receiver; on applying a gentle heat, hydrocyanic acid and watery vapor are disengaged, and condensed in the receiver. The sulphuric acid sets the hydrocyanic acid (which forms part of the ferrocyanic acid in the ferrocyanate) at liberty, by uniting with the potash, with which it was previously in combination. The distillation should be conducted slowly, and stopped when a quantity of liquid, equal to about a third of the water employed, has been condensed; the receiver must be kept cold. The acid, prepared in this manner, often acquires a bluish tint on keeping, from a small quantity of iron, carried over from the ferrocyanate, during distillation, being converted into Prussian blue; it is easily removed by repeating the distillation.

Pure and concentrated hydrocyanic acid is a limpid fluid, like water, but has a strong and penetrating odor, producing severe headache, with nausea, and even fainting, when the vapor which it emits is incautiously inspired. Dr Reid says, he has seen a very stout young man, so much affected by smelling some diluted acid, which had been prepared three months before, from the ferrocyanate of potash, by the process described, that the bottle containing it fell out of his hand, and for half an hour afterwards, he was almost totally unconscious of what was going on around him. It is seldom that the diluted acid affects any one so powerfully, but this shows the great care which should be taken, in making any experiments with this substance. A single drop of the strong acid, introduced into the throat of a large dog, kills it after a few hurried inspirations. Its odor is similar to that of the peach blossom, bitter almond, &c., which indeed derive their peculiar flavor from the presence of a small quantity of this acid. I have seen the diluted acid frequently used a year after it had been prepared, and its strength did not appear to be impaired; but the strong acid is sometimes decomposed in a few hours after it has been made, and can seldom be kept more than a fortnight, at ordinary temperatures.

Its taste is cool at first, but it soon becomes hot and irritating. It evaporates rapidly when exposed to the air; if a drop is suspended at the extremity of a small rod, part of it is congealed by the cold produced by the evaporation of the rest. With

water and alcohol, it combines in all proportions; it reddens the infusion of litmus, slightly, and combines with the salifiable bases, (but does not neutralise them,) for which it has but a feeble affinity, and may be displaced even by the carbonic acid.

As the strong acid boils at the temperature of  $80^{\circ}$ , it may be easily procured in the gaseous form, when disengaged in the first process mentioned for its preparation, by receiving it in pneumatic jars over the mercurial trough, having previously brought the mercury to that temperature, by putting a red hot poker into it for a short time. Its vapor forms an inflammable mixture with atmospheric air, and detonates when mixed with a proper quantity of oxygen gas; the products of its combustion are water, carbonic acid and nitrogen gases. On referring to its equivalent, by volume, it is evident, that 2 measures (one equivalent) will require  $2\frac{1}{2}$  measures of oxygen (five equivalents) for its complete combustion, four of these combining with the carbon of the cyanogen, and forming two equivalents of carbonic acid, while the other converts the hydrogen into water; the nitrogen gas is left mixed with the carbonic acid.

The specific gravity of the diluted acid varies, according to the quantity of water which is combined with it; and Dr Ure has well remarked, that the density of the liquid is a criterion of greater nicety, than most practitioners will be inclined to appeal to, in estimating its strength, diluted acid of the specific gravity 0.998, containing twice as much real acid as when it is 0.996. The test which he proposed for ascertaining the proportion of real acid in the diluted acid, is the peroxyde of mercury, which is dissolved readily by it when reduced to a fine powder, shaking them together in a vial, and continuing to add small quantities of the peroxyde, till it is no longer dissolved. Every two equivalents of the acid ( $27 \times 2 = 54$ ) dissolve one of the peroxyde, 216, which is exactly four times heavier than the acid required, ( $54 \times 4 = 216$ ). On dividing, therefore, the weight of the peroxyde dissolved, by four, the quotient expresses the quantity of real acid contained in the diluted acid employed. No heat must be applied, as it might expel part of the acid at first in the gaseous form, or cause a larger quantity of the peroxyde to be afterwards dissolved. If an excess of mercury has been taken up, the solution turns paper stained with an infusion of red cabbage, to a green.

If any muriatic acid is mixed with the hydrocyanic acid, it may be detected, by adding ammonia till the liquid is neutralised, and evaporating it to dryness by a heat not exceeding  $212^{\circ}$ . Both acids combine with the ammonia, and the hydrocyanate is dissipated, but the muriate remains, requiring a tem-

perature of  $300^{\circ}$  for its volatilisation. Nitrate of silver cannot be employed, in this instance, though the most delicate test of muriatic acid, as the hydrocyanic acid gives a precipitate with a solution of this salt, which it is not easy to distinguish from that produced by muriatic acid.

When hydrocyanic acid has been given in an over-dose, or administered as a poison, the best and indeed the only antidote, according to Dr Herbst, is the cold affusion. He made a number of experiments on animals, and states, that when the dose of the poison was not sufficient to prove fatal, two affusions of cold water, in general, removed every unpleasant symptom; when a larger dose was given, it was found necessary to repeat it more frequently, and to persevere for a considerable time. The certainty of success depends much on the early employment of the remedy. He tried liquid ammonia, repeatedly, which has been much extolled as an antidote to this poison, but says that it will scarcely ever save life, where a dose sufficient to prove fatal has been given, and the symptoms have continued for some time, though he admits, that where the quantity administered is not able to destroy life, it is of great benefit in mitigating the severity of the symptoms. Another evident objection to the use of ammonia, as Dr Herbst remarks, is, that it excoriates the parts to which it is applied, and when sufficiently diluted to be free from this inconvenience, it is of very little use. The cold water was poured freely over the head and back, and afterward over the whole body.

Orfila very properly remarks, that whatever is to be done in a case of poisoning by hydrocyanic acid, should be attempted immediately, as death very speedily ensues, if a large dose be administered.

Some experiments, recorded in a recent number of the *Annales de Chimie*, render it more than probable, that the only chemical antidote for prussic acid is chlorine; and it is now well understood, that the *American Fire King*, as he is called, who swallows strong prussic acid with impunity, prevents any bad results, by the previous use of chlorine water. The experiments referred to, prove, conclusively, that the acid was rendered quite harmless, in the case of dogs, when preceded by chlorine, and that this remedy succeeded, when administered after the baneful action of the poison was manifest.

The operation of chlorine, in these cases, is easily explained; although in making the explanation, we have necessarily to anticipate a little. The hydrocyanic acid, as we have seen, is a compound of hydrogen and cyanogen; the chlorine decomposes the acid, combines with the hydrogen to form the vapor



of muriatic acid, while the cyanogen is liberated. The mode of preparing chlorine water, will be noticed in another chapter.

Dr Christison relates some very interesting cases, in which the chlorine was used with great success. A dose large enough to have killed in fifteen minutes, if no remedy had been used, was rendered harmless by the inhalation of the vapor of hot water, containing one-fourth of its bulk of chlorine gas.

Orfila has likewise advocated the use of chlorine, in these cases, in addition to which, he advises the free use of strong coffee and spirits of turpentine.

It should be remembered, that this acid exists in peach leaves, in bitter almonds, &c., and that the beverage called *noyau*, has done harm when drank freely, by reason of the hydrocyanic acid it contains, derived from the kernels of peach stones, and the juice of peach leaves.

It is proper to say, here, that prussic acid is also produced by the decomposition of animal matters, as in making Prussian blue; and whenever carbon, nitrogen, and hydrogen are evolved, simultaneously, it is plain that they may so combine, as to generate this acid.

The tests proposed for detecting prussic acid are, the nitrate of silver, the sulphate of copper, and the prosulphate of iron.

Drop a solution of the nitrate of silver into diluted hydrocyanic acid; a white, curdy precipitate (cyanuret of silver) falls, which cannot be easily distinguished from Exp. the precipitate that muriatic acid, or common salt, gives with the same solution. This test, therefore, is not much relied on.

Add a few drops of a solution of the sulphate of copper to the hydrocyanic acid, and then a slight excess of caustic potash, also in solution. The potash separates the oxyde of copper, forming a blue precipitate which reacts on the hydrocyanic acid, the oxygen of the oxyde combining with Exp. the hydrogen of the acid, while the cyanogen unites with the copper. The cyanuret of copper is recognised by its white color, after muriatic acid has been added to the blue precipitate, taking care to add no more than is necessary to remove the excess of copper thrown down by the potash.

The prosulphate of iron is employed, precisely in the same manner as the sulphate of copper.

This salt is usually blended with some persulphate of iron, which, although unfit for the experiment by itself, does not prejudice the result, if present in small quantity. Some persons speak of the persulphate as the proper test; but if no prosulphate be in the mixture, Prussian blue will not be formed.



Add to the liquid supposed to contain prussic acid, a solution of green vitriol, then throw down the protoxyde of iron by adding a slight excess of pure potash in solution, and after a few minutes exposure to the air, acidulate the whole with muriatic or sulphuric acid, so as to redissolve the precipitate. The Prussian blue color will then appear, if prussic acid be present.

Dr Ure says, that this acid may be detected by the sulphate of iron, when mixed with 10,000 parts of water, and that the sulphate of copper produces a slight milkiness in water, containing only a twenty-thousandth part.

Leuret and Lassaigne affirm, that this poison cannot be detected in the human body, so late as two days after death, because it is speedily decomposed or volatilised. Where it is thought that death has been occasioned by it, they recommend the intestines to be cut into small pieces, and put into a retort, with their contents and some water, adding a small quantity of sulphuric acid and applying a gentle heat, which should not exceed  $212^{\circ}$ . The volatile products are condensed in a receiver kept cold with ice, and tested in the manner we have described. The odor alone is often sufficient to indicate its presence. Hydrocyanic acid should be kept in well stoppered bottles, and entirely excluded from the light, to keep it, as much as possible, from undergoing decomposition.

### *Ferrocyanic Acid.*

When dry animal matters, mixed with half their weight of the carbonate of potash, are exposed to a dull red heat in an iron pot, and stirred constantly till the pasty mass which is produced at first, ceases to give out fetid vapors, a peculiar compound is formed, which may be separated from the rest of the materials by water, and the solution gives large, yellow colored crystals. They are composed of what is usually termed the ferroprussiate or ferrocyanate of potash\* and water, and it is this salt that is used to prepare the different substances from which cyanogen and its other compounds are usually obtained.

The process for preparing it is very offensive, but, on the small scale, a sufficient quantity may be formed from an ounce or two of materials, to show the nature of the salt obtained in solution, when the dry mass is digested in water and filtered. The minutiae of the reaction that takes place in the first stage

\* Called, also, the *triple prussiate of potash*.

of the process, are still far from having been completely explained. The animal matter, which is composed of carbon, oxygen, hydrogen, and nitrogen, is completely decomposed, part of the nitrogen and carbon combining and forming cyanogen, a portion of which unites with the hydrogen and produces hydrocyanic acid, while the rest combines with some metallic iron, and these two compounds uniting with one another, form what is termed ferrocyanic or ferroprussic acid; the carbonic acid of the carbonate is at the same time expelled, and this new acid remaining in combination with the potash, forms the salt which is obtained by solution in water and crystallisation. According to the most recent experiments, it consists of one equivalent of ferrocyanic acid, 108, two of potash, 96, and one of water, 9; it is therefore represented by the number 213. The ferrocyanic acid, again, may be regarded as a compound of two equivalents of hydrocyanic acid and one of the cyanide of iron. The following diagram shows the quantity of cyanogen, iron, and hydrogen in one equivalent, 108, of the ferrocyanic acid, and the manner in which they must be united according to this view.

Composition of ferrocyanic acid.

Hydrogen	1	-----	27 of hydrocyanic acid.
Hydrogen	1	-----	27 of hydrocyanic acid.
Iron	28	-----	
Cyanogen	26	-----	
Cyanogen	26	-----	
Cyanogen	26	-----	54 of cyanide of iron.
	<hr/>		
	108		108
	<hr/>		

There are two processes for preparing this acid. One consists in adding a solution of 58 parts of crystallised tartaric acid, dissolved in alcohol, to 50 of the ferrocyanate of potash, dissolved in as small a quantity of water as possible. The tartaric acid combines with the potash of the ferrocyanate, forming bi-tartrate of potash, which is precipitated in small crystals, not being very soluble in water, and the ferrocyanic acid remains in the liquid which may be separated by filtration.

In the other process, a solution of the hydrosulphuret of barytes is added, in the first place, to a solution of the ferrocyanate of potash, as long as any precipitation takes place, hydrosulphuret of potash remaining in solution, while ferrocyanate of barytes is precipitated. It is washed on a filter with a small quantity of cold water, then dissolved in one hundred parts of

this fluid, and on adding sulphuric acid to precipitate the barytes, taking care to avoid any excess, the ferrocyanic acid is obtained in solution.

This acid may be obtained in small cubic crystals of a yellow color, by allowing its solution to evaporate spontaneously. It has a much greater affinity for the different salifiable bases than hydrocyanic acid, neutralising them completely, and reddening the vegetable blues; it is not volatile nor poisonous, at least, in small doses, and is slowly decomposed when exposed to the light, the iron which it contains acquiring oxygen, and being ultimately converted into Prussian blue. M. Porrett gave the view of its composition which is here given, though many are inclined to doubt if this compound ought to be considered as a distinct acid, and have regarded it as a compound of the protoxyde of iron with an excess of acid. His opinion has been generally adopted, and appears the most probable, the iron accompanying the other elements that are associated with it to the positive pole, where acids are attracted, on decomposing salts by galvanism, and not going along with the other salifiable bases to the negative pole, where metallic oxydes are always found, when they have been separated by its agency from acids with which they have been combined, and where the iron ought have been disengaged, if it had been in combination with the hydrocyanic acid, in the same manner as potash, or any other base, is united with an acid, in a compound salt. It was called, as first, ferruretted *chyazic* acid, from the initial letters of carbon, hydrogen, and azote, (nitrogen,) but ferrocyanic acid is the term that is now adopted.\*

### *Sulphocyanic Acid.*

This is a compound of sulphur, 32, (or two equivalents,) and hydrocyanic acid, 27, (or one equivalent). For the reason given in the note to the remarks on ferrocyanic acid, I should prefer to call the compound now under consideration, *sulphuretted hydrocyanic* acid. M. Porrett, who is its reputed discoverer, (year 1808,) described it under the name of *sulphuretted chyazic* acid.

This acid is obtained, in combination with water, by adding sulphuric acid to a solution of the sulphocyanate of potash, and

\* I am not disposed to make innovations in nomenclature, but it seems to me, that the more correct name would be *ferrhydrocyanic acid*; this would be a more exact expression of its actual composition, viz. iron, hydrogen, and cyanogen.

distilling the liquid in a glass retort. The solution of the sulphocyanate is prepared by a process introduced by Grotthus, which Dr Turner recommends to be conducted in the following manner. A mixture of equal weights of the ferrocyanate of potash and sulphur in fine powder, is exposed to a strong heat over a charcoal chauffer, taking care, however, not to allow it to become red hot. It soon melts and takes fire, and must be withdrawn from the chauffer a few minutes after the combustion ceases. The residuum consists of sulphur, sulphocyanuret of potassium, and sulphuret of iron; a pure solution of the sulphocyanate of potash is obtained, on digesting it in water, the potassium combining with the oxygen of a portion of water, which is decomposed, while the hydrogen goes to the sulphocyanogen, and the sulphur and sulphuret of iron, mixed with it, remain undissolved; the latter is formed by the sulphur combining with the metallic iron of the ferrocyanic acid, and of the ferrocyanate. When the sulphuric acid is added to this solution, it combines with the potash, and sulphate of potash remains in the retort after the sulphocyanic acid has been separated.

When its solution in water is concentrated, it boils at  $216^{\circ}$ , and crystallises at  $54^{\circ}$ , in six-sided prisms. It reddens the vegetable blues, neutralises the alkalis, and forms a soluble salt, of a deep red color, with the peroxyde of iron. When boiled with iron filings, sulphuret of iron is formed, and hydrocyanic acid disengaged.

#### SECTION VIII.—OTHER COMPOUNDS OF CARBON, HYDROGEN, AND OXYGEN.

I have already alluded to the arrangements of the common treatises on chemistry, which place the consideration of vegetable bodies, far in the rear. A similar defect is noticed, in relation to courses of lectures on chemistry, and hence much valuable information is kept from the student, altogether, or so long deferred, as to be regarded as the least interesting of all topics.

The plan adopted in this work, necessarily gives to vegetable bodies a conspicuous place, because it introduces them precisely in accordance with their obvious relations. We have been reviewing many of the compounds of carbon with other substances, and as the vegetable acids are made up, in great measure, of oxygen, hydrogen, and carbon, our examination would be very defective, if these were not introduced at this time.

Several of these acids, although existing largely in the vegetable kingdom, may be formed artificially, as the oxalic and



hydrocyanic acids; they are not, therefore, always classed along with the rest, though it was at one time considered impossible to produce any thing by a simple physical action, similar to any of the proximate principles formed by an organised structure. The greater number of them can be obtained in the solid form; they are, in general, soluble in water and alcohol, have the same general properties as the mineral acids, but are more feeble in their action, and are all decomposed by a red heat, the carbon and hydrogen which they contain, taking fire if they are exposed freely, at the same time, to the air. When heated in close vessels, the principal products are carbonic acid and carbonic oxyde gases, (composed of carbon and oxygen,) water, acetic acid, and an empyreumatic oil. By nitric acid they are, in general, converted into oxalic acid, and when heated with sulphuric acid, they take away a portion of its oxygen, and sulphurous acid is evolved.

The following table shows the composition of some of the more important vegetable acids. Most of them contain more oxygen than is necessary to convert their hydrogen into water, a circumstance first pointed out by Gay Lussac and Thenard. The oxalic acid contains no hydrogen, and the succinic and acetic acids are generally supposed to contain the same weights of their respective elements. Mr Prout, however, has ascertained, by accurate analysis, that acetic acid contains 3 equivalents of hydrogen, while succinic acid has but 2.

	No. of equivalents.			Carb.	By weight.			Equivalent of the acids.					
	Oxy.	Hydro.			Oxy.	Hydro.	Carb.						
Acetic acid	3	+	3	+	4	=	24	+	3	+	24	=	51
Tartaric acid	5	+	2	+	4	=	40	+	2	+	24	=	66
Citric acid	4	+	2	+	4	=	32	+	2	+	24	=	58
Oxalic acid	3	+	0	+	2	=	24	+	0	+	12	=	36
Benzoic acid	3	+	6	+	15	=	24	+	6	+	90	=	120
Gallic acid	5	+	5	+	9	=	40	+	5	+	54	=	99*
Succinic acid	3	+	2	+	4	=	24	+	2	+	24	=	50
Malic acid	4	+	10	+	3	=	32	+	10	+	18	=	60

The vegetable acids, and, indeed, all vegetable substances, have some marked points of difference from mineral acids and mineral bodies, in general. It is not possible, or rather no attempt has yet succeeded, to form a vegetable compound, by the union of its elementary parts. We are able, however, to reproduce mineral acids, by the combination of their constituents. All vegetable substances being naturally resolved into

\* These numbers correspond, very closely, with the analysis of Berzelius. See *Annals of Philosophy*, Vol. V.



carbon, oxygen, and hydrogen, it occurs, that the vegetable acids are liable to change into each other; of this fact, we have frequent examples. The mineral acids, on the contrary, maintain their distinctive character, and it is not possible to change nitric into sulphuric acid.

### *Acetic Acid.*

From the table presented above, it is apparent that the equivalent number of this acid is 51. This number indicates the absolute acid,\* or, in other words, the acid itself, uncombined with any diluent, and as it exists in connexion with a salifiable base. The strongest acid ordinarily met with, is composed of one equivalent of real acid and one of water.

This acid is known to exist in small quantities, in the sap of many plants, but is usually prepared, by exposing liquids that have experienced the vinous fermentation, freely to the air, at a temperature of about 80°. Many vegetable infusions may be made to pass directly into the acetous fermentation, without the intervention of the vinous, by adding a portion of the fermentable matter, taken from a liquid that has already been subjected to that sort of intestine motion. Thus, cider will soon take on the acetous fermentation, if some of the dregs of a vinegar cask be added. In this process, a large quantity of oxygen is absorbed, and the vinegar which is produced, consists of acetic acid, water, mucilaginous matters that have not been decomposed, and a little alcohol. Large quantities of acetic acid are now obtained, also, by the destructive distillation of wood; this distillation is effected in iron retorts, or cylinders, and yields the *pyroligneous* acid of the shops, meaning, literally, the acid from burnt wood. It is chiefly composed of acetic acid, and empyreumatic oil, and may be purified by a second distillation; after which, lime is added to detach the acetic acid, thus forming acetate of lime. This compound is then decomposed by means of sulphuric acid, which forms a sulphate of lime, and pure acetic acid is then easily distilled.

To prepare strong acetic acid, more directly, pour sulphuric acid carefully, and in small quantities at a time, on twice its weight of the acetate of potash, in a glass retort, waiting till the ebullition ceases after each successive addition of the acid, and distil to dryness with a heat gradually increased, but never very strong, condensing the product in a receiver, surrounded with ice or very cold water. Instead of the acetate of potash, acetate of soda, lime, or lead may be employed, taking care to

\* Formerly called *radical vinegar*.

use rather less sulphuric acid than is necessary to combine with the salifiable base, in the acetate employed; the quantity required may be easily found, by referring to a table of equivalents. In the present process, which is the best, every forty parts of real sulphuric acid, (one equivalent,) which are contained in 49 of the liquid acid, decompose one equivalent of the acetate, 98, consisting of 48 parts of potash and 50 of acetic acid, combining with the potash, and forming a sulphate which remains in the retort, while the acetic acid and water of the sulphuric acid are distilled over. It always contains a little sulphurous acid, from which it may be separated by a second distillation, mixing it with a little of the acetate of lead.

The Edinburgh college directs it to be prepared, by putting six parts of the dry sulphate of iron with five parts of the acetate of lead into a retort, after they have been well mixed, and distilling from a sand bath till no more acid is disengaged, condensing the product in a receiver. Here, the sulphuric acid of the sulphate unites with the oxyde of lead in the acetate, and disengages the acetic acid, oxyde of iron and sulphate of lead remaining in the retort.

Strong acetic acid may be procured, also, by exposing some of the metallic acetates to heat, without any addition. To prepare a little in this manner, fill a small green glass retort half full of the acetate of copper, and expose it to heat over a good charcoal chauffer, condensing the product in the usual way; if a green glass retort cannot be obtained, take one of flint glass, and heat it by a sand bath, or coat it with plaster of paris, and then the chauffer may be used. The acid obtained in this manner is never pure, being always mixed with pyroacetic spirit, formed by a part of it being decomposed.

Acetic acid and water unite in all proportions, and the specific gravity of the compound is sometimes the same, though very different quantities of acid and water may be combined together. The following table shows its specific gravity, when combined with one, four, and seven atoms of water. (Dr Thomson.)

Specific gravity at 60.	Acid.	Water.
1.06296	1 equivalent	1 equivalent
1.07132	1 _____	4 _____
1.06349	1 _____	7 _____

Its density, therefore, cannot be taken as an index of its strength.

The weak acetic acid of the different colleges is prepared by

distilling vinegar, rejecting the first part out of every eight or ten, as it contains very little acid; the next five or six constitute the weak acetic acid of the pharmacopœias, and the distillation is stopped when they have come over, or the product collected in a different receiver, as it then has an empyreumatic odor, from the mucilaginous matters which it contains, beginning to be decomposed. Its specific gravity varies from 1.006 to 1.009; 1000 grains, of the latter specific gravity, are neutralised, according to Mr Phillips, by 145 grains of the crystallised carbonate of soda. It cannot be rendered stronger by distillation alone, as the acid and the water rise together on the application of heat. By exposing vinegar or a weak acid to cold, most of the water freezes, leaving a stronger acid which may be separated by straining.

The method of estimating the strength of acetic acid is, by ascertaining the quantity of crystallised carbonate of soda, which it can neutralise; 50 parts (one equivalent) of real acid being required for one equivalent of the carbonate.

Strong acetic acid has a very pungent and agreeable odor, and volatilises rapidly when exposed to a moderate heat, producing an inflammable vapor, which is easily kindled. It crystallises at a low temperature, and remains solid till heated again to 50°. Its acid properties are very well marked, neutralising completely the different salifiable bases, and forming salts that are decomposed by heat, and easily distinguished by the odor of acetic acid, which they emit when sulphuric acid is poured upon them; it reddens the vegetable blues, powerfully, oxydates iron, copper, lead, zinc, and some other metals, and raises a blister on the skin, when kept in contact with it for some time. It dissolves volatile oils, camphor, and the active principles of some of the most powerful vegetable medicines. The smelling salts that are sold in the shops, are sulphate of potash, mixed with this acid; or rather, these articles constitute one species of those salts.

Acetic acid is often sold in a very impure state. Sulphuric and sulphurous acids may be detected by acetate of lead, which gives a white precipitate when they are present. Copper will render it blue, on adding an excess of ammonia, and lead may be detected by sulphuretted hydrogen, which will give a black precipitate.

Acetic acid is used, in the shape of vinegar, to extract the medicinal virtues of squill and colchicum; it has also been employed in the formation of the *black drop*. In diseases of the urinary organs, attended with white deposits, or *sand* in the urine, acetic acid has been given to effect the decomposition of

the phosphate of lime, or the phosphate of magnesia and ammonia, of which those deposits chiefly consist.

The impure pyroligneous acid has been found to act as an antiseptic. It not only prevents putrefaction, but stops it after the process has commenced. These effects have been ascribed to the peculiar empyreumatic oil which the acid contains; and as it cures and preserves herring, hams, tongues, &c., equally well with the practice of smoking, it has been inferred, that the active principle of the smoke and the acid, is the same. The peculiar flavor imparted, is alike in both cases. One quart of pyroligneous acid, added to a barrel of hams, in the ordinary pickle, will give them the flavor of smoke, and cure them completely. Indeed there is no necessity for using a pickle; if the hams are dry salted, and then hung up and washed four or five times with the acid, they will be well cured.

It has generally been supposed, that acetic acid was not a poisonous substance; but Orfila has shown,\* that, in its concentrated form, it is destructive to animal life. Its antidotes are the same articles that are proper to counteract the mineral acids.

### *Of Tartaric Acid.*

This is sometimes called the acid of tartar. It is composed of 40 oxygen, (five equivalents,) 2 hydrogen, (two equivalents,) and 24 carbon, (four equivalents). The crystallised acid contains one equivalent of the water of crystallisation, viz. 9, making the equivalents of the dry and crystallised acids to be as 66 to 75. It is soluble in five parts of water, at 60°.

This acid may be formed by pouring 49 parts, by weight, of sulphuric acid, on 94 of tartrate of lime, diffused in four times its weight of boiling water, stirring the mixture occasionally, for a day or two, after it has been well rubbed in a mortar, and evaporating the liquid, obtained by filtration, through a linen bag. One equivalent of the tartrate of lime, 94, is composed of 66 of acid and 28 of lime, and the 40 of dry sulphuric acid, in the quantity employed, combines with the lime, and forms 68 of the sulphate of lime, while the tartaric acid is disengaged, and remains in solution. Very little of the sulphate of lime is dissolved, as it requires a large quantity of water for its solution, and the tartaric acid crystallises when the liquid is evaporated, which should be done in an earthen evaporating bason, over a sand bath, with a very moderate heat. The sulphate of lime, in the liquid, is deposited after the evapora-

\* Journal de Chimie Medicale, August, 1831.



tion has been continued for some time, and should be separated, by pouring it into another vessel; when it has assumed a syrupy consistence, it may be set aside to crystallise. The crystals must be purified by a second crystallisation.

The tartrate of lime is prepared, by mixing chalk, (carbonate of lime,) in fine powder, intimately, with four times its weight of cream of tartar, (composed of 180 bi-tartrate of potash + 18 water,) and throwing the mixture into 10 or 12 times its weight of boiling water, adding small quantities at a time, that the effervescence which takes place may not be too violent; 94 parts of solid tartrate of lime are precipitated, 114 parts of neutral tartrate of potash remain in solution, and 22 of carbonic acid are disengaged. The following diagram gives a view of the reaction that takes place.

Before decomposition.		After decomposition.	
180 bi-tartrate of potash, or 1 equiv.	{ potash 48 tartaric ac. 66 tartaric ac. 66	114 tartrate of potash.	
50 carb. lime, or 1 equiv.	{ carbonic ac. 22 lime 28	22 carbonic acid.	
		94 tartrate of lime.	
230	230	230	

An additional quantity of tartrate of lime may be obtained, by adding a solution of the muriate of lime to the solution of the tartrate of potash, as long as any precipitation takes place, if it is not required for other experiments. The proportions in which the materials react on each other, are represented below.

Before decomposition.		After decomposition.	
114 tart. of potash, or one equiv.	{ potash 48 tartar. ac. 66	85 muriate of potash.	
65 muri. of lime, or one equiv.	{ muri. ac. 37 lime 28	94 tartrate of lime.	
179	179	179	

Tartaric acid crystallises in prisms; its solution in water is very sour, but has an agreeable taste when sufficiently diluted; it is also soluble in alcohol. When exposed to heat, it is completely decomposed, and by destructive distillation, a peculiar acid is formed, which has been termed, from its mode of preparation, the *pyrotartaric acid*. Tartaric acid is particularly



distinguished by forming an insoluble salt with potash, (bi-tartrate of potash,) when added in excess to this alkali, though the neutral tartrate of potash and the tartrate and bi-tartrate of soda are very soluble

Tartaric acid and many of its salts have the property of dissolving many metallic oxydes, and of preventing them from being precipitated from their solutions, by substances which are in general capable of producing this effect. This may be easily seen, by adding a little tartaric acid to a solution of the sulphate of iron, and then some ammonia in excess, which will not precipitate the oxyde of iron, after it has been mixed with the tartaric acid.

Tartaric acid decomposes carbonates with effervescence, and precipitates potash from its solutions, when they are not diluted with water, small crystals of the bi-tartrate being formed.

The solutions of tartaric acid become mouldy, by long keeping, as do the solutions of the other vegetable acids. This change amounts to an entire decomposition, and the ultimate elements, oxygen, carbon, and hydrogen, enter into new relations, and form other secondary compounds.

Tartaric acid may be substituted for citric acid, (of which we are next to speak,) in the formation of refrigerant drinks. It does not deliquesce by exposure to the air, and is usually employed in the formation of the effervescent powders, which are employed in the place of soda water. For this purpose, the acid should be powdered, then dried by a gentle heat, and kept in well stoppered bottles. A small quantity is to be dissolved in three or four ounces of water, and about as much bi-carbonate of potash or soda in another vessel, containing the same quantity of water. The two solutions are then poured into one vessel, and effervescence immediately ensues. Or, the acid and salt, well powdered, may be kept mixed in a tight vessel; a teaspoonful of this mixture, stirred into a small tumbler nearly full of cold water, makes a pleasant effervescing draught, and presents an excellent vehicle for the administration of Epsom salts.

Seventy-six parts of crystallised tartaric acid, will saturate 70 of dry carbonate, and 101 of crystallised bi-carbonate of potash. This fact determines the proportions in which the acid and salt should be mixed, to form the effervescing draught.

In addition to the mode for obtaining tartaric acid, already noticed, it should be known, that this acid exists in many vegetables, and that it has been manufactured very advantageously, from the juice of sour grapes; thirty-six ounces of which, yield two and a half ounces of the acid.

If tartaric acid be heated in close vessels, it gives out carbonic acid and carburetted hydrogen gases. If nitric acid be poured on the crystals of tartaric acid, and the whole be distilled, the tartaric will thus be changed into oxalic acid.

### *Of Citric Acid.*

The name of this acid is derived from the fact, that it is made from the juice of lemons or limes; or rather because it actually exists in those fruits. Its equivalent number and ultimate composition are given in the table, appended to the general remarks on vegetable acids. But the number 58, is there given as the equivalent of the *dry* acid; the *crystallised* acid contains 18 (or 2 equivalents) of water, which makes its actual equivalent 76. It is soluble in less than its own weight of water, at 60°, but boiling water dissolves twice as much. It is also soluble in alcohol.

In order to prepare citric acid, proceed as follows. Take a given quantity of lemon or lime juice, and add finely powdered chalk, as long as any effervescence occurs. The chalk, being a carbonate of lime, is decomposed, its carbonic acid escapes, causing the effervescence, while the citric acid, of the lemon juice, unites with the lime, and forms a citrate of lime. This compound may be detached from all mucilaginous matters, by frequent washing with water; and one equivalent, or 86 parts, are decomposed, completely, by 49 parts (or one equivalent) of sulphuric acid. The latter, taking the lime from the citrate, forms a sulphate of lime, while the citric acid remains in solution. By filtering and evaporating the acid solution, crystals of citric acid are obtained, which require, in order to be made quite pure, a second solution, filtration, and crystallisation. The following diagram exhibits the decomposition, effected by the action of sulphuric acid on the citrate of lime.

49 or 1 equivalent of common sul- phuric acid	{	water	9	—	9 or 1 equiv. water.
	{	real sulphu- ric acid	40		
86 or 1 equivalent citrate of lime	{	citric acid	58	—	58 or 1 equiv. citric acid.
	{	lime	28	—	68 or 1 equiv. of sulph. of lime.
135			135		135

Citric acid bears a great resemblance to tartaric acid, but may be easily distinguished from it, by carbonate of potash,

with which its solution gives no precipitate, when added in excess. Tartaric acid may be detected in this manner, if it has been mixed with citric acid, a practice that is occasionally followed, as it is much cheaper than the citric acid. A solution of the same strength as lemon juice is obtained, according to Mr Phillips, by dissolving nine and a half drachms of citric acid in a pint of water. The specific gravity of the lemon juice varies considerably, and according to Dr Henry, a wine gallon usually affords from six to eight ounces (avoirdupois) of the crystallised acid.

This acid is employed in medicine to form effervescing mixtures, in all which cases it effects decomposition. It is added to some one of the carbonates, and while it unites to the bases of those salts, it expels the carbonic acid, and hence, the effervescence. It is owing to the free evolution of carbonic acid gas in these cases, that such mixtures often relieve nausea and arrest vomiting. Twenty grains of bi-carbonate of potash are saturated by 10 grains of crystallised citric acid; the same weight of the carbonate requires 15 grains. Dr Paris gives the following convenient table, to show what quantities of citric acid and lemon juice should be employed, to neutralise the alkalis.

Citric acid.	Lemon juice.	A scruple of alkali.
Grains x	3 fluid dr'ms.	Bi-carb. potash
Grains xv	4 do. do.	Carb. of potash
Grains xxv	7 do. do.	Carb. ammonia

The *incompatibles* of citric acid are, necessarily, all the alkaline and earthy carbonates; to which may be added, the alkaline and metallic acetates, the sulphurets of earths and alkalis, and alkaline soaps.

The juices of many other fruits besides the lemon, contain citric acid, as the cranberry, red gooseberry, wild strawberry, &c.; in all cases, it is, in common with the other free vegetable acids, a proximate principle. On account of its ready solubility, this acid is highly useful to manufacturers. The calico printers employ it to *discharge* colors, as they say; an operation that is mainly chemical.

#### *Oxalic Acid.*

This has been denominated *acid of sugar*, *acid of wool*, *sorrel*, &c. The name *oxalic* is derived from the *oxalis acetosella*, from the juice of which it has been obtained, although it exists in several other plants.

The equivalent of the dry acid is 36. In the crystallised acid there are four equivalents of water, 36, making the whole equivalent 72. It is soluble in thrice its weight of cold water, and in its own weight of boiling water. It is also soluble in alcohol.

Oxalic acid may be formed by the action of nitric acid on sugar, starch, gum, wool, and other substances that contain the elementary principles which are common to animal and vegetable matter, especially carbon, hydrogen, and oxygen. It is most frequently made of sugar.

To illustrate the process by which it is obtained, fill a flask or retort, about a third full of nitric acid, (an ounce or two will be a sufficient quantity on the small scale,) and add a sixth part of its weight of refined sugar, coarsely powdered, in small quantities at a time. Apply a gentle heat, if necessary, by a lamp or chauffer, to commence the action; let the heat then be removed till it becomes feeble, evaporating the remaining liquid, till it acquires the consistence of syrup, when it may be set aside to crystallise. Every 100 parts of sugar give about 60 of crystallised oxalic acid; it must be purified by solution in water and a second crystallisation.

The precise nature of the reaction which takes place has not been very minutely investigated; small quantities of malic and acetic acids are produced at the same time, and, as oxalic acid contains no hydrogen, and the quantity formed is much less than the weight of sugar employed, its formation must depend on the oxygen of the acid attracting the hydrogen and part of the carbon from the saccharine matter. A large quantity of nitrous acid, nitric oxyde, and carbonic acid is disengaged; the fumes are very offensive, and the flask should be placed where they may be carried off.

Oxalic acid crystallises in four and six sided prisms, has an extremely sour taste, and reddens sensibly the vegetable blues, even when dissolved in 3000 parts of water. It is resolved into carbonic acid and carbonic oxyde, when exposed to heat. It gives a white precipitate with lime water, forming an insoluble oxalate of lime.

Oxalic acid is a powerful poison, and has proved fatal, occasionally, even when diluted with a large quantity of water, and taken as an acidulous drink; two or three drachms are sufficient to produce death. Drs Christison and Coindet, in an able memoir on poisoning by oxalic acid, in the Edinburgh Medical and Surgical Journal, have shown that chalk and magnesia are certain antidotes to this poison, when administered in proper time, the oxalates of lime and magnesia, which are formed, being quite inert.



Most cases of poisoning, by this acid, have arisen from its having been mistaken for Epsom salts, (sulphate of magnesia,) to which it bears a considerable resemblance in external appearance, when the crystals are small; they are easily distinguished from each other, however, by a number of simple tests. Oxalic acid has a very sour taste, reddens the vegetable blues, effervesces with a solution of the carbonates of potash or soda, and is completely dissipated by heat; sulphate of magnesia, on the contrary, has a pure, bitter taste, does not affect the vegetable colors, gives a white precipitate, with solutions of the alkaline carbonates mentioned, but does not produce any effervescence, and parts with its water of crystallisation, on exposure to heat, without undergoing any further change. An emetic should be taken, followed by large draughts of warm water to promote its action, in the first instance, or the stomach pump may be employed, at once. Some persons have advised emetics to be given instantly, but if the patient be vomiting, they need not be administered, nor should we wait for the stomach pump, if the antidotes are at hand. Much less should warm water be given with a view to accelerate vomiting, for the dilution of the poison will only increase the surface of its action.

Formerly, the alkalis were resorted to, but they are altogether improper, since the oxalates, thus formed, are almost as fatal as the acid itself.

The practitioner should be exceedingly careful to label his bottle of Epsom salts very distinctly, so that he may not confound this article with the oxalic acid, if the latter be in his shop. Apothecaries have made fatal mistakes, for want of this precaution, as the acid and the salt are frequently so much alike, in general appearance, that they cannot be distinguished by the eye.

All the oxalates are decomposable by heat. Those that are soluble in water, give copious precipitates with salts of lime, and if the oxalates, thus precipitated, are digested in sulphuric acid, sulphate of lime will be formed, while the oxalic acid will be held in solution.

### *Of Benzoic Acid.*

The name of this acid is derived from the gum benzoin, from which it is obtained. It contains a larger relative quantity of carbon, than any other vegetable substance of which we have any knowledge. Its composition, as seen in the table, is 90 carbon, (15 equivalents,) 24 oxygen, (3 equivalents,) and 6 hydrogen, (6 equivalents). It melts and begins to be volatilised at 230°. It is dissolved in 24 parts of boiling water, but is more soluble in alcohol.



A convenient mode of preparing this acid, is to take one or two ounces of the gum, in powder, and place it in a glass or earthen vessel, and cover it with a tall, paper cone. The whole is then placed in a sand bath, which is to be gradually heated, until white fumes begin to rise and to fasten on the sides of the cone. I generally use a common sand bath on a close stove, and have the cone so large, at its base, as to surround the vessel containing the gum. When the cone is fixed in its place, the sand is raised around its base, so as to prevent the escape of the acid vapors.

Another process for obtaining this acid is preferred by many, as it is not then contaminated with any empyreumatic oil, which is always mixed with it when prepared by sublimation. The Edinburgh college directs three parts of the gum to be reduced to a fine powder, and intimately mixed with one part of the carbonate of soda. The mixture is then boiled for half an hour, in twenty-four parts of water, the liquor poured off, and the residuum boiled again with nine parts of water. The benzoic acid combines with the soda of the carbonate, forming benzoate of soda, which remains in solution, and the carbonic acid is separated with effervescence; the other parts of the benzoïn are not dissolved. The mixed decoctions are then filtered and evaporated, till only two parts remain, and on adding sulphuric acid, previously diluted with seven parts of water, as long as any precipitation takes place, the benzoic acid is thrown down, and sulphate of soda remains in the liquid, from which it is separated by filtration. It is afterwards dried and sublimed, to obtain it in the light, feathery, and crystalline form, in which it is known by the name of *flowers of benzoïn*. Both processes may be easily conducted with an ounce or two of the gum.

Benzoic acid exists in several other productions of the vegetable world, and in the urine of children and graminivorous animals.

Pure benzoic acid is very white, and has a shining lustre. Its odor is fragrant and peculiar, but M. Giese attributes this to the presence of a small portion of oil. It burns with a yellow flame when exposed suddenly to a strong heat.

Benzoic acid has the property of separating the peroxyde of iron completely from its neutral solutions, when combined with soda or ammonia, and is used frequently for this purpose in analytical chemistry, especially when the iron is associated with manganese, which it does not precipitate.

The only pharmaceutical preparation into which benzoic acid enters, is the *tinctura opii camphorata*, usually called *paregoric elixir*. The salts formed by this acid, are termed *benzoates*.

*Of Gallic Acid.*

As this acid is found most abundantly in nutgalls, it has derived its name from that source. The galls, being well bruised, are macerated in water; then the liquor is filtered, and permitted to stand exposed to the air, for some two or three months. During this period, it becomes mouldy, is covered with a thick, glutinous pellicle, flocculent particles fall to the bottom, and the sides of the vessel are covered with small yellowish crystals, which may also be found on the under surface of the supernatant pellicle. These crystals may be purified by solution, and evaporated to dryness.

M. Deyeux advises, to put the powdered galls into a glass retort, and to apply heat very cautiously. The acid rises and is condensed in the neck of the retort. But this process is so very uncertain, that few attempt it.

M. Baruel forms gallic acid, by adding the solution of white of egg to the infusion of nutgalls, till it ceases to have any obvious effect. The clarified solution is then evaporated to dryness, and the residuum dissolved in alcohol. This new solution is next filtered and concentrated by heat, until it is reduced to the proper state for yielding the acid, by crystallisation.

It is said, that gallic acid, in whatever way obtained, contains a portion of another acid, that has been called *ellagic acid*.

Gallic acid is found in different vegetable matters, that have astringent properties, as well as in the nutgall. It is soluble in water and alcohol. It takes fire when exposed to heat, and is distinguished by the dark blue precipitate it gives with salts of iron, and which constitutes the basis of black ink. When combined with tannin, (a substance to be spoken of hereafter,) with which it is usually associated, it has the important property of precipitating most of the metallic oxydes from their solutions, even when combined with the more powerful acids, and hence its use as a reagent. For this purpose, an infusion of gall-nuts is generally used; it is made by pouring a few ounces of boiling water over an ounce of galls, and allowing the mixture to stand for an hour or two, before filtration. The tincture keeps better, but has no other advantage over the infusion; it is prepared by digesting the powdered galls in alcohol.

Prepare a vial of pure water, and drop some tincture or infusion of galls into it; also, a vial with a clear solution of sulphate of iron, (green copperas,) and mix these clear fluids in one vessel. As soon as they come in contact, the whole is quite black.

Write with a weak solution of sulphate of iron, and when dry, the writing will be invisible. Dip a feather in the infusion or tincture of galls, and draw it over the writing, and presently the letters will be black and legible. Exp.

From these experiments, it is easy to imagine, that gallic acid is an important agent in the process of dyeing; it is largely employed, though not in its separate state, in all establishments for printing calicoes.

Gallic acid is composed of carbon, oxygen, and hydrogen, as may be seen by the table before referred to. Its salts are called *gallates*.

### *Of Succinic Acid.*

This name comes from the Latin word for amber, and the compound is often called the *acid of amber*.

Succinic acid may be formed, by exposing powdered amber, mixed with an equal weight of sand, to heat, in a green glass or coated flint glass retort. The retort must be heated by a good charcoal chauffer or spirit lamp. Some oily matter first comes over, and then succinic acid is deposited, in crystals, on the neck of the retort and sides of the receiver. These are to be purified by dissolving them in a watery solution of potash, and boiling the solution with charcoal. The potash combines with the succinic acid, forming succinate of potash; to this, nitrate of lead is added, and a double decomposition occurs, giving rise to succinate of lead and nitrate of potash, the latter remaining in solution, and being easily separated by filtration. To the precipitate of succinate of lead, we next add sulphuric acid, which combines with the lead, and leaves the succinic acid free. The sulphate of lead is perfectly insoluble, and, of course, can be easily separated from the acid.

Succinic acid has a sour taste, and reddens litmus paper. It is soluble in water and alcohol, and the solutions are easily crystallised by evaporation. Its salts are called *succinates*.

### *Of Malic Acid.*

The name of this acid is derived from the Latin word for apple. As the table shows, it is a compound of four equivalents of oxygen, 32, ten equivalents of hydrogen, 10, and three equivalents of carbon, 18, making its entire equivalent 60.

It is found in unripe apples, and in the juices of several other fruits. The acid obtained from the berries of the *sorbus acuparia* was formerly called *sorbic acid*, but it is now known to be

the same with malic acid. Scheele is the individual who made the discovery of this acid, and he announced it in the year 1785.

Many processes have been given for the separation of malic acid from its connexion with vegetable substances. In all cases, it is requisite to add lime, lead, or some other base, to form a *malate*, which is easily decomposed by sulphuric acid; an insoluble sulphate being formed, while the malic acid is set free.

This acid, as well as its compounds, are not of sufficient importance to claim much notice in an elementary work; and the reader is referred to the fuller treatises, for further and more particular information.

**OF GUM.** We mean by this term, the mucilage of vegetables. It is usually transparent, more or less brittle when dry, and not easily reduced to powder. It has an insipid, or slightly saccharine taste, is soluble in water, giving it a gluey or adhesive consistence, which varies according to the proportion of gum in solution. It may be separated, or coagulated by the action of weak acids; it is insoluble in alcohol and in oils, and is capable of the acetous fermentation, when diluted with water.

The best specimen of a pure gum, is the *gum acacia* or *gum arabic*, being that which is so frequently used in medicine. The specific gravity of this gum is 1.515, and the purest sort is almost white, and very transparent.

If four ounces of powdered gum arabic, be gradually rubbed with a half pint of boiling water, there will be formed a mixture of the consistence of syrup, to which the term *mucilage* has been applied. This article is much employed in the preparation of pharmaceutical compounds, and is admirably calculated to effect the incorporation of many substances that could not otherwise be united without difficulty.

Gum is also soluble in pure alkalis and lime water, as well as the vegetable acids, (especially vinegar,) with which it forms a mucilage, that may be used as a cement, without the risk of becoming mouldy. This latter quality is owing to the action of the acid employed; for the hangers of paper always add solutions of alum to their paste, to prevent it from becoming mouldy, and they are thus successful.

Gum is insoluble in alcohol, and hence all infusions and decoctions that hold gum in solution, give precipitates when alcohol is added, if the gum be in sufficient quantity. A very simple

Exp. experiment proves this position. Dissolve some gum arabic in water, and pour the solution into alcohol; the water will seize the alcohol, and let the gum fall in a dense, opaque precipitate.

Gum arabic is demulcent and nutritious, and has been the entire diet of patients, for several successive weeks. In some conditions of the body, however, it passes the bowels unchanged.

*Gum Senegal* is sometimes sold for *gum arabic*, but it has a clammy and tenacious quality, whereas gum arabic is dry and brittle.

If we add a solution of oxalic acid to a watery solution of pure gum, the acid will unite with the small portion of lime which gum always contains, forming an oxalate of lime, which appears in form of a white precipitate. Exp.

Add a solution of gum arabic to a solution of the subacetate of lead; the gum is instantly precipitated in combination with part of the oxyde of lead. Gum is thrown down from its solution by some other salts, as the persulphate of iron, and pernitrate of mercury. Exp.

Let a solution of silicated potash be added to a solution of gum arabic;\* a precipitate will soon appear, which is regarded as a compound of the lime contained in the gum and silica. The pure alkalis form soluble compounds with gum. Exp.

The stronger acids, when considerably diluted, have no peculiar action on gum arabic, except by long standing, further than to cause its solution. If we mix some of the gum, in fine powder, with strong sulphuric or nitric acid, it will be completely decomposed. The sulphuric acid will cause a deposition of charcoal, and the nitric acid will change it into oxalic acid, if a large quantity be employed.

The composition of gum is ascertained to be the same, from whatever source obtained, viz. oxygen, carbon, and hydrogen.

Besides the oxalic acid, obtained from gum by the action of nitric acid, we have, at the same time, portions of malic and *mucic* or *saccholactic acids*. The two last are one and the same compound, and may be readily formed by digesting gum in three times its weight of nitric acid. On applying a lamp heat, effervescence ensues, and the acids are formed.

The *mucic* or *saccholactic acid* was discovered by Scheele in 1780. Its formation depends on a new arrangement of the component parts of gum, with an accession of oxygen, derived from nitric acid. Thus, while in 100 parts of gum there are but fifty parts of oxygen, we find 61 parts of the latter in 100 of *mucic acid*.

\* These solutions may be made, by adding one or two ounces of good gum arabic to a pint of water, a little warm.



	Carbon.	Oxygen.	Hydrogen.
100 parts gum contain	42.23	50.84	6.93
100 parts mucic acid	33.6	61.5	4.9

**SUGAR.** This is an abundant vegetable product, and as it exists in many ripe fruits, and in other vegetable matters, it is often called a proximate principle of vegetables. In this country, the maple yields sugar in large quantities; so also does the sugar cane. In France, it is largely obtained from the beet, and the sugar is of good quality.

As the sugar, in the maple and ripe cane, exists in the sap, in a diluted form, the principle of extraction depends on the separation of the excess of aqueous particles, by boiling, evaporation, and the like. The use of lime water in the process, is to neutralise any free acid that may be present, and also to promote the separation of extractive and other vegetable matters. When the juice is sufficiently reduced by boiling, it is drawn off into shallow vessels, kept quite cool, and there it soon appears in form of loose, crystalline grains.

The same principle may be illustrated, by dissolving some sugar in hot water, till a strong syrup is obtained, and then setting it aside to crystallise. In a few hours, crystals will be deposited, resembling, in appearance, small portions of sugar candy. If we dissolve the sugar in alcohol and set the mixture aside, we obtain large crystals; four parts of hot alcohol, take up about one of sugar.

Molasses, sugar of grapes, honey, and manna, are only varieties of the saccharine matter, which is so widely diffused throughout the vegetable kingdom.

There has been some disagreement about the elementary nature of sugar; but it appears now to be agreed, that its equivalent is 15, being composed of carbon 6, oxygen 8, and hydrogen 1. To this constitution of sugar, we had occasion to refer, when speaking of the formation of alcohol.

Sugar rapidly decomposes nitric acid, especially if heat be employed. Large quantities of dense fumes of nitrous acid are evolved, and the elements of the sugar are so deranged and recombined, as to form oxalic acid. Sugar, however, contains a portion of hydrogen, which is not found in the oxalic acid resulting from this process.

	Carbon.	Oxygen.	Hydrogen.
Oxalic acid contains	12	24	0
Sugar	6	8	1

Like tartaric acid, sugar has the property of preventing oxide of iron and some other metallic oxides from being precipitated from their solutions, by some of the reagents that usually

throw them down; it possesses this property, however, only when boiled with a solution of the metallic salt. Sugar partially deoxydates several of the metallic oxydes, when their solutions are heated together.

Wood, starch, linen, glue, and many other vegetable and animal substances may be converted into sugar, by the action of sulphuric acid. This results from the fact, that all these substances contain the elements of sugar, and it may therefore readily occur, that these may be so disposed of, as to give rise to the saccharine product. It is usual to mix the animal or vegetable matter with about twice its weight of sulphuric acid, diluting the mixture in a day or two with a large quantity of water, and boiling it for six hours or more, adding water from time to time, in small quantities, to supply that which is lost by ebullition. The liquid is then to be neutralised with powdered chalk, and filtered; then the clear liquid is evaporated, till it attains a syrupy consistence, after which it may be set aside to crystallise.\*

**STARCH OR FECULA.** This is another very abundant vegetable matter. It exists in almost all the grains, in several roots, as the potato, and in the kernels of leguminous plants. A very simple method of procuring it, is to let a small current of water fall upon the dough of wheat flour, enclosed in a piece of linen, subjecting it at the same time to pressure between the fingers, until the liquid passes off quite clear. The gluten of the flour is left in a pure state, the saccharine and mucilaginous matters are dissolved, and the starch is washed away, mechanically, being deposited from the water on standing, in the form of a white powder. The starch of commerce is made by an analogous process, and large quantities of it are procured from the potato.

We distinguish starch from gum and sugar, by its insolubility in cold water, and by forming a jelly with hot water. It is insoluble in alcohol or ether. The aqueous solution is precipitated by subacetate of lead, which is decomposed, as are several other metallic salts; with silicated potash, it gives no precipitate. The best test of starch, is a substance to be spoken of hereafter, and called *iodine*.

\* Recently, I examined some white sugar, suspected to contain arsenic. The color of the common *ratsbane* is very much like that of white sugar, and may be easily mixed with it. But water, of the temperature of common *house-tea*, while it takes up the sugar, does not dissolve the arsenious acid, and, in this way, the poisonous attempt may be detected. In the instance referred to, the design was palpable, but fatal results did not follow, for want of a sufficient dose of the poison. The metal was reduced, in the way already described. See *arsenic*.

The very useful article of diet, called *arrow-root* or *Indian arrow-root*, which is prepared from the root of the *maranta arundinacea*, has all the characteristics of pure starch. *Sago* and *tapioca*, both derived from East Indian plants, are, chemically, the same substance. From these facts, it may be seen why the potato starch may be substituted in the room of arrow-root and tapioca, (as we know it often is,) without injury to our patients. *Salep*, which is yielded by the *orchis mascula*, contains, with other things, a portion of starch.

One hundred parts of starch contain 43.55 carbon, 49.68 oxygen, and 6.77 hydrogen, which are nearly the same proportions that are found in sugar. This close resemblance in the constitution of starch and sugar may serve to explain the fact, that the former is convertible into the latter. This change is evident in the germination of seeds, as in the malting of barley, in which process, the starch of that grain passes into the saccharine state. This result is effected by a suitable degree of heat; but we find that the abstraction of heat operates in the same way, as is seen in the agency of severe frost on potatoes, apples, and parsnips.

The *mucilago amyli*, mucilage of starch, has been much employed by some practitioners of medicine. It is made by triturating three drachms of starch with a pint of water, and then gently boiling until a mucilage is formed. This preparation has been chiefly exhibited with opium, in form of clyster, in diarrhœas and dysenteries, where the tenesmus arises from abrasion of the mucous coat of the rectum.

**GLUTEN.** This has been referred to in the remarks on starch. It may be made, by forming wheat flour into a stiff paste with water, and washing it in a linen cloth or bag, with a large quantity of this fluid, as long as any white powder passes through the interstices of the cloth. A very convenient mode is to hold the dough under the rapid stream of a hydrant, when that can be had. The gum and sugar are dissolved by the water, and carried off; the starch is removed in form of white powder, while the gluten, only, remains in the cloth. It is a very tenacious and ductile mass, insoluble in water, and when moist, is easily decomposed, owing to reaction of its elements. It is regarded as the most nutritious part of wheat flour, and in addition to carbon, oxygen, and hydrogen, the peculiar elements of vegetable matter, it is found to contain nitrogen. This is one of the sources of the nitrogen of animal bodies, in general, and seems to be a sort of connecting link between vegetable and animal life.

Gluten is of a gray color, and fibrous structure, exceedingly

viscid and elastic. If it be dried by a gentle heat, its volume is much diminished, and it may then be preserved for years, without change.

Mr Taddei, an Italian chemist, has obtained from gluten two distinct principles, which he calls *gliadine*, from the Greek word for *gluten*, and *zymome*, from the Greek word for *ferment*. The gluten is to be rubbed with successive portions of alcohol, in a mortar, so long as this fluid becomes milky on the addition of water; the alcohol dissolves the gliadine and leaves the zymome.

The gliadine, obtained from the alcoholic solution by evaporation, is a brittle, slightly transparent substance, of a yellow color, and sweetish, balsamic taste. Its smell, when cold, resembles that of the honeycomb; when hot, its odor is like to that of boiled apples. It is soluble in boiling alcohol, but is deposited on cooling. Water makes the alcoholic solution quite milky, and the gliadine is thrown down. It is insoluble in water, but is dissolved by acids and alkalis. Heated in open air, it takes fire and burns with a bright flame.

Zymome is a hard, tough substance, without the viscosity of gluten. It is insoluble in water and alcohol; but soluble in vinegar and the mineral acids, with the aid of heat. With pure potash, it forms a soap. It is susceptible of putrefaction, without previously fermenting, as gluten does; and when heated, it gives out an odor like that of burning hair. It occasions various kinds of fermentation, according to the nature of the substance with which it comes in contact.

A very delicate test of zymome has been announced by Mr Taddei. He found, that on mixing the fine powder of gum guaiacum with zymome, a beautiful blue color was formed, especially if made a little moist; and the same result is obtained, though less rapidly, if gluten or wheat flour be rubbed with guaiacum and moistened with water. The intensity of the blue color depends on the relative quantity of zymome in the flour, and this, on the quantity of gluten; the proportion of the latter, and thence the quality of the flour, may be estimated, with tolerable accuracy, by the action of guaiacum.

Now the cause or precise nature of the change which gives rise to the blue color, has not been explained. Of one thing, however, we are certain, and that is, the agency of oxygen gas in the phenomenon; for if atmospheric air be wholly excluded, the change does not follow.

**TAN OR TANNIN.** This substance exists in a great number of vegetable bodies, more especially in those that possess astringency; and it is, usually, associated with gallic acid. It may

be obtained in a pure state, by precipitation, from an infusion of nutgalls, by means of the permuriate of tin. The precipitate consists of tannin and oxyde of tin, and must be diffused in a large quantity of water, after which the metallic oxyde must be separated, by the action of a stream of sulphuretted hydrogen gas. Thus, sulphuret of tin is formed, while the tannin is held in solution.

*Artificial tannin* has been formed by Mr Hatchett, by digesting charcoal and some other vegetable matters, in nitric acid; in the opinion of some chemists, however, the product is merely a compound of nitric acid and charcoal.

Tannin is distinguished by its astringent taste, by precipitating a number of metallic oxydes from their solutions, and by forming a copious precipitate with solutions containing gelatine, (glue,) similar in composition to leather, which is, in reality, no other than a compound of tannin and gelatine. In the process of tanning leather, the hides are cleaned of their hair by the caustic operation of lime, after which the tannin of the bark combines with the gelatine of the hide, and leather is the product.

Tannin is very soluble in alcohol, and in hot water, and is precipitated from its solutions, by acids. In its dry state, it is a brown, friable substance, and of a resinous fracture. Its aqueous solution has a deep brown color, and it does not become mouldy by keeping.

To a thin solution of glue, add a small quantity of tannin, either alone or in form of a strong aqueous solution, and a yellowish, flocculent precipitate subsides, which is insoluble  
 Exp. in water, resists putrefaction, and, on drying, becomes hard and tough. Tannin is, therefore, a test for gelatine.

If tannin be added to a solution of sulphate of iron,  
 Exp. (green copperas,) a black color will result. It is on this principle that our black ink is formed, of which more will be said, when we treat of the salts of iron.

One hundred parts of tannin contain 45 oxygen, 50.55 carbon, and 4.45 hydrogen.

**VEGETABLE ALKALIS OR ALKALOIDS.** By the term *alkali*, we mean a body, which, besides its disposition to unite with acids, and form salts that are neutral in their character, has the power of changing the purple colors of vegetables to green, the reds to purple, and the yellows to a brown. If the purple be reddened by an acid, the alkalis restore the color. These are the principal features, which are important to be stated here. The subject will be more fully noticed, when we treat of potash, &c.



The term *alkaloid* has been preferred by some authors, inasmuch as the vegetable alkalis, so called, are supposed to want some of the marked characteristics of true alkalis. But as in several respects they resemble alkalis, the term *alkaloid* has been selected, because it indicates no more than a resemblance.

I have hesitated, more than once, respecting the location of this important branch of vegetable chemistry. But when I recollect, that custom has almost uniformly postponed it to the close of a course of lectures, and, frequently, by the press of other matter, it has been quite excluded, I feel warranted in placing it in that position to which it appears to be entitled, by natural affinity. Many of our pages have been already occupied with the compounds of carbon, and the other simple, non-metallic substances that have been examined; but as all the articles usually brought into view under the denomination of *vegetable* alkalis, are kindred, in their elementary constitution, with many that have just passed in review, there seems to be a fitness in the examination of those bodies in this place.

The limits of this work will not permit a full range through the entire field that now lies before us. It must suffice to notice a part only, referring the reader, for details, to the larger treatises. Among the more important, we select the following, viz. *morphia*, *quinia*, *cinchonia*, *strychnia*, *brucia*, and *veratria*, all of which are composed of carbon, hydrogen, oxygen, and nitrogen. The other proximate principles that have acquired some celebrity, are all composed of the same elements, in varied proportions, and their extraction from the vegetable matters in which they exist, is conducted on the same general principles. What may be said of the preparation of the articles above-named, will admit of practical application to most, if not all the vegetable alkalis, now known.

The honor of discovering this class of bodies, is due to Ser-tuerner, a German apothecary, who published an account of *morphia* in the year 1803; the subject excited no particular notice, however, until the publication of his second essay in 1816. The chemists who have since cultivated this important field, most successfully, are Robiquet, Pelletier, and Caventou.

All the vegetable alkalis are decomposed by heated nitric acid, and ammonia is always a product of the destructive distillation. They never exist in an insulated state in the plants that contain them, but are, probably, in every case, combined with an acid, with which they form salts, more or less soluble in water. These (pure) alkalis are, for the most part, insoluble in water, and sparingly soluble in cold alcohol; but they are all readily dissolved by that fluid at a boiling temperature, and

when the solution is cold, they are deposited at the bottom, and on the sides of the vessel, in form of crystals. The salts of these alkalis are, generally, more soluble than the alkalis themselves, and some of them are remarkable for their solubility.

The outline of the processes that are usually resorted to, is as follows. The substance, in which the alkaline principle resides, is digested or macerated in a large quantity of water, which dissolves the salt, the base of which is the pure, vegetable alkali. On adding some more powerful, salifiable base, such as potash or ammonia, or boiling the solution for a few minutes with lime or pure magnesia, the vegetable alkali is separated from its acid, and falls as an insoluble powder. It may then be collected, by throwing the whole on a filter, and washing with water. As thus obtained, it is not quite pure, but is combined with some oleaginous, resinous, or coloring matters, with which it is associated in the plant. To purify it from these, some ivory black (animal charcoal) is added, and the whole is then boiled in alcohol. As hot alcohol dissolves these alkalis, they are thus taken up; and if filtered while hot, and then set aside to evaporate and cool, the pure alkali will remain, while the alcohol is dissipated. The necessity of using much alcohol is sometimes obviated by the following plan. The alkali, after having been precipitated and collected on a filter, is combined with some acid, such as the acetic, sulphuric, or muriatic, and the solution boiled with animal charcoal, until the coloring matter is removed. The alkali is then precipitated from its connexion with the acid, by ammonia or some other base, which readily unites with the acid, and so takes the place of the vegetable alkali.

I have no doubt that all the vegetable alkalis will yet be obtained independently of alcohol; and this will not only simplify the process, but tend to diminish the use of this article in the laboratory, as well as to lessen its importance in society.

**OF MORPHIA.** This term is derived from *Morpheus*, the god of sleep. Morphia exists in opium, and is the active, efficient principle whose operation induces sleep. Besides this vegetable alkali, opium contains a peculiar vegetable acid, viz. the *meconic*, and a third vegetable, proximate principle, destitute of acid as well as alkaline properties, and called *narcotine*. These substances are intimately blended with several more ordinary vegetable bodies, as gum, resin, gluten, a bitter matter, a dark coloring matter, and also a small portion of sulphate of lime.

Sertuerner, long ago, noticed the deposition of morphia, in vessels containing the alcoholic solution of opium, or the laudanum of the shops, and Dr Hare having recently made a state-

ment of a similar fact, expresses his astonishment, that morphia was not investigated at a much earlier period. But it must be remembered, that the attention of scientific men was not awake to these subjects, until within a few years; and it would seem that the spirit of research is now sufficiently active to insure the exploration of the entire vegetable kingdom, until every proximate principle shall be brought to light.

Several processes are in use for the extraction of morphia from opium. I shall notice but two, of which Robiquet's is the first. Take 1200 grains of the best Turkey opium, cut it into small pieces, and rub it with water in a mortar, adding a little more, from time to time, till a uniform, pulpy mass is obtained. Then add more water till about ten ounces, in all, shall have been mixed with the opium. The mixture is to be left in this state five days, shaking it frequently, or it may be digested with a gentle heat, (not over  $100^{\circ}$  Fahrenheit,) for a day or two, after which it may be filtered. The liquid that passes through the filter has a very dark color, and contains the meconic acid and the morphia in solution, combined in form of meconate of morphia. On boiling this solution for ten or twelve minutes with 60 grains of magnesia, (entirely free of carbonic acid,) the meconic acid will dissolve part of the magnesia, and remain in solution, as meconate of magnesia, while the morphia will be precipitated. It is still mixed, however, with some coloring matter, which may be removed by washing it with cold water, until the latter passes unchanged; afterwards with a little hot and cold alcohol alternately, till no more coloring matter is taken up, and on boiling the residuum in an ounce or two of alcohol till it is dissolved, crystals of morphia will be deposited from the solution as it cools, in a great measure free from coloring matter.

This direction, however, is not essential in all its parts, to the formation of good and active morphia. The frequent washings with water will make the article sufficiently pure for practical uses, without the necessity of resorting to alcohol. To get rid entirely of the coloring matter, some of the processes direct the use of animal charcoal, and it has a very obvious effect in this way.

A process is on record, in a volume of the Journal of Science, by means of which its contrivers (Messrs Henry and Plisson) procured good morphia, independently of the use of alcohol. It is founded on the fact of the facility with which morphia may be separated from opium, by means of weak muriatic acid. The muriate of morphia is decomposed by ammonia, and the morphia obtained in a pure form by the aid of charcoal. If it

is desirable to have it in handsome crystals, it may be dissolved in alcohol and evaporated.

The *Journal de Chimie Medicale*, &c., for February, 1830, contains a process for preparing morphia, by the fermentation of opium. Two pounds and three fourths of the purest opium were reduced to small pieces, placed in a vessel with a large aperture, and covered with twice its weight of lukewarm water, in which three ounces of honey and nearly two ounces of yeast were diffused. The mixture was then placed in a stove, moderately heated, and fermentation soon commenced. When the intestine motion had ceased, the mixture exhaled a very distinct smell of alcohol, indicating something like the vinous fermentation. The whole was then placed in a tight bag, and pressed very forcibly, fresh water having been, in the meanwhile, poured several times upon it. The liquor thus forced through the bag, after it has become cool, is treated with an excess of ammonia. By this means, a precipitate is formed, which must be washed with cold water and dried. This dried precipitate is next to be reduced to powder, and water poured on it that has been slightly acidulated with muriatic acid. The liquid is now of a yellowish-brown color, and after the acid has been acting for some hours, the liquor is to be filtered, and evaporated to the solid state. The muriate of morphia, thus procured, is highly colored, but after it has been well washed with pure water and boiled with animal charcoal, it appears in the form of silky crystals, of a beautiful pearl white. It is from this muriate of morphia, that the pure morphia is obtained, by pouring into its aqueous solution, a very slight excess of liquid ammonia. The ammonia takes the muriatic acid from the morphia, while the latter is thrown down, in form of a marble colored powder. It is next to be carefully dried, and the crystals, on exposure to the light, exhibit a very brilliant appearance.

This morphia is said, by Mr Blondeau, the inventor of the process, to be of excellent quality; and we are assured, that a larger quantity is procured in this way, than by any other method in use. It need only be added, here, that a committee of the Royal Medical Society of Paris, confirm the statements of Mr Blondeau.

In whatever way morphia is prepared, care should be observed by the operator, in relation to the vapors that are evolved, especially if a large quantity of materials be employed; for it not unfrequently happens, that the mere fumes produce the same effects that result from an overdose of opium.

It is affirmed, that lime is preferable to magnesia, for the



purpose of precipitating morphia from its connexion with meconic acid, as it naturally exists in opium. The quantity of precipitate is declared to be larger, and of purer quality.

The power of opium to induce sleep, seems to reside immediately in the morphia, the taste of which is intensely bitter. Both the pure alkali and its salts are supposed by many practitioners to possess peculiar advantages, which are not to be obtained from opium alone; and I confess myself to be of that number, although aware that some respectable physicians think otherwise.

Orfila has affirmed, that pure morphia is nearly inert, owing to its comparative insolubility; but I believe he is incorrect, and such is the general impression. It does not act so efficiently as some of its salts, and this is doubtless owing to the circumstance already referred to, viz. its want of solubility.

To show that morphia has alkaline properties, dip a test-paper, colored with the infusion of blue cabbage, into an alcoholic solution, and it will be changed to green, instantly. Put another test-paper, colored by turmeric, Exp. into the same solution, and it will be changed to a reddish-brown.

The compound usually called *sulphate of morphia*, may be easily formed by the union of morphia and dilute sulphuric acid. These are to be shaken together and gently heated, morphia being added in small quantities until no more can be taken up. The acid is, in this way, completely neutralised, and a true sulphate is formed. It is exceedingly soluble in water, and may be very advantageously dissolved in mint water.

The acetic acid may be made to combine with morphia in a similar manner, and then we have the *acetate of morphia*. This acid is active in the formation of the well known article, *black drop*, and there can be no doubt, that acetate of morphia is formed in that mixture. The laudanum of Rousseau, is also indebted, for its energetic powers, to the formation of the same salt, as it is evident that acetic acid is generated in the process. So also the *liquor opii sedativus*, a popular anodyne in London, is said to owe its efficacy to the acetate of morphia.

The *nitrate*, *muriate*, *citrate*, *tartrate*, and *carbonate of morphia*, are all soluble and crystallisable, and may be formed on the same principle as the sulphate and acetate.

As the vegetable acids readily combine with morphia, and form soluble salts, we see the reason why they should not be exhibited to persons who have taken excessive doses of opiates. But as these acids, especially the acetic and citric, tend to counteract the bad effects of narcotics, they are proper remedies



after the opiate has been ejected by vomiting, or has been removed by the stomach pump. I have seen the happiest effects produced by lemon juice in such cases.

To show the precipitation of morphia from its salts, and so to illustrate the mode of its separation from the meconic acid, with which it is combined in opium, we may resort to the following easy process. Add a solution of potash, soda, or ammonia, to a solution of a salt of morphia. The alkali will combine with the acid, and the morphia will be thrown down. Morphia is also precipitated when a solution of any of these alkalis is added to a watery infusion of opium, a meconate of potash, soda, or ammonia, remaining in solution. With the exception of ammonia, however, none of these alkalis can be so advantageously used for the preparation of morphia as magnesia, because any excess would react on other matters in the solution, and tend to decompose them, or to dissolve the precipitated morphia.

Pure morphia is colorless, or nearly so, yet it is changed to a bright red, as are all its salts, by the action of a very little nitric acid. Put a few grains into a test tube, then pour  
 Exp. on it a little nitric acid, and apply a gentle heat. The mixture will soon acquire a red color, and a portion of oxalic acid be formed. To make the latter apparent, neutralise the acid in the colored mixture, by adding ammonia diluted with three or four times its bulk of water, and drop into it a solution of muriate of lime. The oxalic acid will immediately seize the lime, by virtue of superior affinity, and an oxalate of lime will appear in copious precipitate.

If we add a solution of potash to a solution of sulphate of morphia, the sulphuric acid will quit the morphia to join the potash, and morphia will be precipitated.

Exposed to heat, morphia is entirely decomposed, in the same manner as other vegetable products. It consists of carbon 72, nitrogen 5.50, hydrogen 5.50, and oxygen 17, in 100 parts, and these elements being disturbed by the agency of caloric, enter into new combinations.

M. Scrullas has ascertained, that the compounds of the simple substance, *iodine*, are good tests of the presence of morphia; when the iodic acid, or the acid iodates, are brought into contact with morphia or its salts, iodine is precipitated, in form of a black powder, if the solutions be sufficiently concentrated. This effect will be quite obvious, if a hundredth part of a grain of morphia be present, but the precipitate will then be of a reddish-brown.

As the *meconic acid* and *narcotine* are closely connected with the consideration of *morphia*, a few remarks on each of these substances will be necessary.

### *Of Meconic Acid.*

Several processes have been pointed out for the preparation of *meconic acid*; the following method, by which it is easily obtained, was recommended by Dr Hare. Add a solution of the subacetate of lead to a watery infusion of opium, as long as any precipitation takes place; wash the precipitate on a filter, till the water passes colorless, then diffuse it through a quantity of water, and pass a brisk stream of sulphuretted hydrogen through it for a quarter of an hour or twenty minutes. The precipitate that is thrown down by the subacetate of lead, is composed principally of meconate of lead, which is afterwards decomposed by the sulphuretted hydrogen gas, sulphuret of lead and water being formed, while the disengaged *meconic acid* remains in solution. The sulphuret of lead must be separated by filtration, and the reddish, amber-colored solution that passes the filter, after being exposed to heat, to expel the excess of sulphuretted hydrogen, and to evaporate it, yields crystals of *meconic acid*.

We may also separate *meconic acid* from the meconate of lead, by means of sulphuric acid; in this instance, an insoluble precipitate of sulphate of lead is formed, from which the *meconic acid* may be separated by filtration. An excess of sulphuric acid increases the difficulty of obtaining the *meconic acid* in crystals.

In preparing *morphia* from opium, by means of *magnesia*, we form meconate of *magnesia*, which, by double decomposition with sugar of lead, (acetate of lead,) yields meconate of lead, from which the *meconic acid* may be detached, by either of the processes above named.

*Meconic acid* is soluble in water and alcohol; it has a sour taste, reddens litmus paper, and is particularly distinguished by the dark red color which it produces when added to a solution of a persalt of iron. Drop a small quantity of the watery solution of *meconic acid*, into a glass of water containing a little of the muriate or sulphate of the peroxyde of iron, and a deep red color will be instantly formed. This is so delicate a mode of detecting the presence of *meconic acid*, that (agreeably to Dr Hare) it succeeds very distinctly, when only ten drops of laudanum are diffused in a gallon of water. The mode of procedure adopted by Dr H. was as follows. A

Exp.

few drops of a solution of sugar of lead (acetate) were added to the water containing the laudanum, which was placed in a conical glass vessel, that the precipitate which appeared in several hours, and was deposited on the sides of the vessel, might be easily collected at the bottom, by stirring it gently, from time to time, with a glass rod. The precipitate thus formed, was the meconate of lead, the acetic acid of the acetate having combined with the morphia, forming acetate of morphia which remained in solution. The supernatant liquid, containing the acetate of morphia, was then removed, and 30 drops of sulphuric acid were poured over the meconate of lead, by means of a dropping tube, after which an equal quantity of the solution of persulphate of iron was added in the same way, when the characteristic deep color appeared.

In this process, the sulphuric acid effected the decomposition of the meconate of lead, setting the meconic acid free, which, in union with the chalybeate solution, gave the distinctive color.

The power exhibited by meconic acid, of reddening solutions of iron, was known to Vogel, Robiquet, and others, before the publication of Dr Hare's observations on this subject.

Robiquet directs meconic acid to be separated by the following method, by which, he affirms the largest quantity can be obtained. He treats opium with magnesia, and so forms meconate of magnesia. This is decomposed by means of muriate of barytes, and two compounds are the result, viz. muriate of magnesia, and meconate of barytes. To the latter, dilute sulphuric acid is added, for the purpose of detaching the barytes and setting the meconic acid free.

**OF NARCOTINE.** This is a word of Greek derivation, meaning an agent capable of inducing stupor. It is a white, insipid, inodorous, crystalline substance, exerting no action on vegetable blues. It is soluble in alcohol, ether, and oils, but insoluble in water. If a small portion of acid be added to water, it will dissolve a considerable quantity of narcotine. In the watery infusion of opium, it is probably kept in solution, by the meconic or some other vegetable acid.

Some persons have attributed the unpleasant sensations, so often produced by opium, to narcotine; and as this principle is soluble in ether, it may be separated from opium by that agent. On this point, however, there are different opinions, and some late writers have not hesitated to declare, that the *denarcotised opium*, or laudanum, is not, in any sense, a better article than the pure opium itself. It is alleged that narcotine occasions stupors and convulsions, but does not induce proper sleep; and some individuals make the same objections to opium, in partic-

ular cases. Dr Tully, of Yale College, affirms that laudanum deprived of narcotine, (denarcotised laudanum,) differs in no respect from common laudanum, except in being considerably weaker, and in having the flavor of ether. He gives this as the result of his own experience, and that of his friends.

When opium, or any of its preparations, has been given as a poison, it may, in general, be detected by its peculiar odor, and by the deep red color struck by the addition of a solution of persalt of iron. After having filtered the contents of the stomach, acetate of lead is added, which gives rise to a meconate of lead. On separating this precipitate, and decomposing it, by means of a few drops of sulphuric acid, a very little of the solution of persulphate of iron, will strike a deep red color, thus indicating the presence of meconic acid, which is always contained in opium.

We know of no certain antidote for the poison of opium, or its preparations, that can act on chemical principles. It has indeed been affirmed, that chlorine and iodine will decompose all the vegetable alkalis, by taking their hydrogen, and so changing their very nature; but we are not yet in possession of a sufficient amount of evidence on this subject, to warrant us in coming to a favorable conclusion. The carbonate of potash has been proposed, in cases of poisoning by the salts of morphia, in order to precipitate the morphia from its solutions, and thus diminish the poisonous action; but I am not inclined to give much credit to this suggestion. It does not appear to be established, that morphia, uncombined, is inert; and we know that acids are present in the stomach, with which it may readily unite, and so give rise to highly active compounds.

The stomach pump, powerful emetics, and every contrivance by which the stomach may be emptied of its contents, should be resorted to, but acids are entirely improper, while any portion of the noxious matter is in the primæ viæ; although they are useful after the complete evacuation of the offending matter, by restoring the tone of the stomach and bowels, and improving the whole system.

**OF QUINA OR QUININA.** This article is usually called, in this country at least, by the name of *quinine*. It is the most valuable vegetable alkali that has been discovered, and will probably retain its high character, while febrile diseases attack the human system. When the disagreeable taste and often nauseating action of the large doses of Peruvian bark, are recollected, it is matter of astonishment, that human ingenuity had not sooner ascertained the active principle of that medicine. If a few grains of quinine contain all the febrifuge virtues of an



ounce of bark, it is manifest that our stomachs have been compelled, for years, and with much labor, to accomplish a result which is now among the common operations of the laboratory. How greatly are we indebted to chemistry for thus coming to our relief, and presenting us with a remedy that scarcely has a parallel in the entire range of *Materia Medica*!

The quinine exists in several species of bark, but most largely in the yellow and pale bark, the former of which is now generally preferred.

Quinine may be easily obtained by decomposing a solution of the sulphate of quinine in water, by means of slaked lime. The lime combines with the sulphuric acid of the sulphate of quinine, and an insoluble sulphate of lime is thrown down, and the quinine, being insoluble in water, is also precipitated. As the quinine is soluble in alcohol, a portion of that liquid is added, which takes up the quinine and leaves the sulphate of lime unchanged. If the alcoholic solution be evaporated, the quinine will remain in the vessel, in form of powder.

We show the alkaline character of quinine, by putting a blue test-paper into an alcoholic solution; it is instantly changed to green. Turmeric paper, in like manner, is altered to a reddish-brown color. If we add quinine to sulphuric acid, diluted in the usual way, we are able to form a neutral mixture, which contains sulphate of quinine, and which may be separated, in fine crystals, by evaporation.

As quinine is composed of the ordinary elements of vegetable matter, it yields, on being burnt, the usual products of the combustion of vegetable matter, as water, carbonic acid gas, &c.

Owing to the insolubility of pure quinine in water, it is less useful as a medicinal agent, than the various salts of which it is the base. Of these, we have several, as the *sulphate*, *acetate*, and *phosphate*, all of which are employed in the treatment of fevers.

**SULPHATE OF QUININE.** This is the most important salt of quinine, possessing all the virtues of the pure alkali, and being sufficiently soluble in water to enable it to act more powerfully and uniformly. It has an intensely bitter taste, similar to that of quinine.

Several methods are in use for the preparation of the sulphate, of which the following is probably the best. Boil two (troy) pounds of yellow bark (in powder) in two wine gallons of water, previously mixed with two ounce measures of oil of vitriol; strain the decoction through a cloth of linen, and boil the residue again, with a fresh quantity of soured water, and refilter. To the decoctions mixed together, gradually add powdered lime, until the liquid has become slightly alkaline, and of a dark



color, which usually requires about half a pound of lime. A brown, flaky sediment falls down, which is separated by straining through a linen cloth, washed with a little cold water, and then exposed to a gentle heat. When this sediment is dry, it is to be digested in several successive portions of spirit of wine, with a moderate heat, for some hours, until all the bitterness is extracted. The several portions of spirit are then mixed and distilled with a gentle heat, until three-fourths of the spirit have passed over. The residue in the body or matrass, is a brown, thick substance, covered with a bitter alkaline liquid, which is to be poured off, saturated with weak sulphuric acid, and boiled down with a little ivory black. The liquor is then filtered while hot; on cooling, the sulphate of quinine crystallises, and the crystals are to be dried on blotting paper.

The brown, thick substance is again boiled in a small quantity of water, slightly soured with the oil of vitriol, which changes a considerable portion of it into sulphate of quinine.

Two pounds of yellow bark, generally yield from five to six (apothecaries) drachms of the sulphate of quinine, in crystals of a pearly lustre.

In this process, the sulphate of quinine is first formed, combined with various impurities; the lime added decomposes this sulphate, while sulphate of lime is formed, and the quinine, comparatively pure, is liberated. The sulphuric acid, then added, unites to the quinine, and forms the sulphate of quinine.

This is a very bitter, white salt, susceptible of crystallisation, in fine, pearly needles, which accumulate in light masses. Its solubility in water, is increased very much by the addition of a little sulphuric acid, which converts it into a super or bi-salt; hence the practice of adding to solutions of this salt in water, a few drops of the elixir of vitriol, which is, in fact, a dilute sulphuric acid.

Messrs Henry and Plisson have published a mode of preparing quinine, without the use of alcohol, and it cannot be doubted that all the important vegetable alkalis may be prepared by similar means. They substitute fresh hydrate of lead for lime, adding it to the decoction of bark, until the sulphuric acid is a little more than neutralised, and has a faint, yellowish tint. The kinic acid of the bark combines with the lead, and forms kinate of lead, which exists in the yellowish liquid, in union with the kinates of lime and quinine. Some coloring matter combined with oxyde of lead, sulphate of lead, and free quinine are deposited at the bottom of the vessel. The quinine is easily changed to sulphate of quinine, by sulphuric acid properly diluted.

Henry has also given the following short method for making the sulphate. He treats bark several times with water acidulated with sulphuric acid, in proportion of a half ounce to a quart of water, applying a moderate degree of heat. The liquor is then to be filtered through linen cloth, treated with quick-lime, and the precipitate washed with water, to remove any excess of lime. This being well drained, is repeatedly digested with strong alcohol. The spirituous tinctures being collected, are placed over a water bath. Then the alcohol is distilled off, to be used in future operations. The residue is a brown, viscid, bitter substance, which contains impure quinine. This mass is exposed to heat, with some diluted sulphuric acid, and the liquid, when cool, yields crystals of sulphate of quinine. A second solution and crystallisation renders the crystals quite pure.

Sulphate of quinine is capable of assuming a phosphorescent appearance. To exhibit this property, place some of the salt in a glass bottle, and introduce it, cautiously, into hot water.

As sulphate of quinine has always commanded a high price, it has been the subject of frequent adulteration. Earthy matters are detected by exposing it to a red heat, when all the sulphate of quinine will be decomposed or volatilised, leaving nothing but the earthy matter behind.

It has been adulterated with sugar, and also with wheat flour. To detect the addition of sugar, dissolve the suspected sulphate in water, and add a solution of the carbonate of potash, which will immediately precipitate the quinine. Filter the liquid, evaporate, and digest in alcohol. If any sugar be present, it will be dissolved by this liquid, and on evaporation, the sugar, only, will remain.

If flour be suspected as the adulterating agent, the powder will be made viscid and pasty, by wetting it with water.

Recently, one of the component parts of most fatty matters, viz. *stearine*, has been employed to adulterate the sulphate of quinine. It very closely resembles the salt in external appearance, but may be easily separated. If the mixture be acidulated with sulphuric acid, a sulphate of quinine, with excess of acid, or a bi-sulphate will be formed, which is very soluble in water; the stearine, being insoluble, is detached, and may be melted into a greasy fluid, by applying a gentle heat.

The sulphates of morphia and quinine closely resemble each other, in external appearance. If to solutions of these salts we

Exp. add a few drops of nitric acid, we shall find the solution of morphia changed to a deep red, while the quinine will be unaltered.

Quinine is susceptible of combination with acetic and phosphoric acids, and the resulting compounds are the *acetate* and *phosphate* of quinine. The latter has been very highly recommended in the treatment of fevers, by Dr Harless, of Boun, in Germany. He affirms it to be more agreeable to the palate than the sulphate, and that it is better suited to the purpose of animalisation, by reason of the animal acid it contains. It is also said to be admirably adapted to irritable stomachs, and vastly superior to any other preparation of quinine, in such cases. The dose is from one to four grains.

Oxalic, tartaric, and gallic acids combine with quinine, and form salts. For this end, the oxalate, tartrate, and gallate of potash, in solution, may be mixed with solutions of sulphate of quinine; double decomposition follows, and the oxalate, tartrate, and gallate of quinine are produced. They are all insoluble in cold water, but dissolve in hot water or alcohol.

**OF CINCHONIA OR CINCHONINE.** This vegetable alkali is prepared from the pale bark, in the same manner in which quinine is procured from the yellow bark, and the two alkalis closely resemble each other in their leading properties. Cinchonine differs from quinine, in being easily crystallised, in the quantity of acid which it neutralises, and in some minor circumstances. It is composed of the same elementary substances, joined in different proportions.

**KINIC ACID.** This acid is combined with cinchonine in the bark, and as it has been referred to, it is proper to say a word or two respecting its formation. It is united to lime, as the kinate of lime, which may be procured by the spontaneous evaporation of the liquid obtained by the maceration of bark in water. If oxalic acid, in solution, be added to the kinate of lime, decomposition is effected, oxalate of lime is formed, and the kinic acid, being thus detached, may be obtained in crystals, by evaporation.

Before we quit the subject of bark, entirely, it may be well to observe, that chemistry furnishes a convenient mode to distinguish between good and bad varieties of bark. If an infusion of galls be poured into a decoction of bark, a copious precipitate will appear; the gallic acid which it contains, uniting with the alkali of the bark, and forming gallate of quinine or cinchonine. But if the alkali has been extracted from the bark previous to making the decoction, or if the bark be of bad quality, little or no precipitate will be thrown down.

**STRYCHNIA.** The celebrated Java poison and St Ignatius's bean, owe their active powers to the presence of the peculiar alkali called strychnia or strychnine. This is a salifiable,

vegetable base; a white, pulverulent substance, having an excessively bitter taste. If examined with a microscope, it will be found to be composed of many small, prismatic crystals. It is scarcely soluble in boiling water and ether, but dissolves well in alcohol and the volatile oils. Exposed to the action of fire, it is decomposed, giving off ammonia. It unites with acids, forming salts that are soluble in water and crystallisable. These salts are sometimes neutral, sometimes acid, and sometimes with an excess of base. It is obtained, directly, by double decomposition. The strychnine is found in different parts of the *strychnos*, in combination with an acid; it is usually extracted from the *strychnos nux vomica*, which also contains a quantity of another vegetable alkali, viz. *brucine*. An infusion of the vomica nuts is treated with magnesia; the precipitate is then treated with alcohol, which takes up the greater part of the brucine; after this it is treated by concentrated boiling alcohol, which unites with the strychnine and deposits it on cooling; it is purified by recrystallisation. It is composed of carbon 78.22, nitrogen 8.92, hydrogen 6.54, and oxygen 6.38. It was discovered in 1818, by Pelletier and Caventou.

Although this article has obtained considerable reputation as a remedial agent, it is, like many other active agents, a virulent poison, when taken in large doses. It induces all the usual symptoms of tetanus, and finally, in a short time, destroys life. It is said that chlorine and iodine are its best antidotes, as the compounds, thus formed, are inert. The chlorine probably destroys the poison of the alkali, by detaching its hydrogen.

Three grains of strychnine, dissolved in one ounce of alcohol, make a tincture of proper strength; of this, from 6 to 24 drops are a suitable dose. It is also used in form of pill and watery solution.

**BRUCINE.** Brucine or brucia, is an alkaline substance, of a pearly white color, similar to boracic acid. It is sometimes in a spongy mass, little soluble in water, insoluble in ether and fixed oils, but soluble in hot and cold alcohol; submitted to the action of fire, brucine is decomposed, and ammonia is formed. Brucine combines with acids, and salts are the result; some of these are neutral, and some have the acid in excess. It is very bitter, notwithstanding its insolubility. Pelletier and Caventou discovered it in 1819, in the bark of the *brucea anti dysenterica*, where it is united to gallic acid. It has also been found in the *strychnos ignatia* or *Ignatius's bean*, as we have already hinted. The *brucea* yields it better. The process is as follows. Make a strong decoction of the plant with water, and then add oxalic acid, which takes the brucine from the gallic acid, with which it



was united. Then evaporate the liquor to an extractive consistence; the residue is treated with cold alcohol, which dissolves all the matters, except the oxalate of brucine. This salt is then boiled with an equal part of magnesia, by which it is decomposed, an oxalate of magnesia formed, and brucine liberated. The free brucine is then redissolved in alcohol, and filtered, in order to separate the oxalate of magnesia; by a slow evaporation, the alcohol deposits minute, regular crystals. This substance is very poisonous, but less so than the strychnine. It is distinguished by the property of becoming of a blood red, when a few drops of strong nitric acid are poured upon it. It is not employed in the practice of medicine, but being less active than strychnine, there is reason to believe it might answer better.

What has been said of the treatment for an overdose of strychnine, is applicable to brucine. In addition to those remarks, it is proper to say, that emetics are useful, in order to evacuate the stomach; but if the stomach pump is at hand, it ought to be resorted to, without waiting for the operation of emetics. Whatever is attempted, should be done without delay.

OF VERATRINE OR VERATRIA. This article was obtained in 1819, from the seeds of the *veratrum album*, or white hellebore, and from the bulbs of the *colchicum autumnale*, or meadow saffron, by Pelletier and Caventou. To procure this alkaline matter, the pounded seeds of the *veratrum* are first treated with sulphuric ether, which dissolves an oily matter, a coloring matter, and an odorous acid; the residue is several times treated with boiling alcohol, the different solutions being cooled, filtered, and evaporated to an extractive consistence; which is to be redissolved in cold water. The aqueous solution, on being evaporated, deposits a reddish matter; when the liquor has acquired a degree of concentration, it is precipitated by the acetate of lead, and then a current of sulphuretted hydrogen gas precipitates the excess of lead, and a colorless liquid is obtained. On treating this with magnesia, the veratrine is precipitated along with the magnesia; the veratrine is then dissolved, by boiling the precipitate in alcohol. On filtering and cooling, the veratrine is, in part, deposited; the other portion is obtained, by evaporating the alcoholic solution. Analysis shows it to be a compound of carbon 66.75, nitrogen 5.04, hydrogen 8.54, oxygen 19.60. According to M. Majendies' experiments, the veratrine produces similar effects on the human system with hellebore, but it is more energetic.

Veratrine is a white, inodorous substance, very sharp to the taste, but not bitter. It fuses at 122°, changing into a white,



waxy mass; at higher temperatures it is decomposed, and ammoniacal products are given out. It is soluble in ether and alcohol, insoluble in cold water, and boiling water scarcely dissolves the 1000th part; yet this small quantity communicates a very sharp taste. It possesses alkaline properties. Strong nitric acid decomposes it, without producing a red color, as it does with brucine. If this acid be added in excess, it very quickly resolves the vegetable matter into its elements, giving birth to a yellow, detonating matter.

As veratrine produces the same effects as the plants whence it is extracted, it may doubtless be given in cases to which they are adapted, due regard being had to its superior energy on the system. Majendie and others assure us, that the veratrine is a better medicine than any of the common preparations of colchicum, and that it may be employed, in all cases, in which colchicum is administered. It is given in pill, powder, and tincture. A half grain, mixed with syrup and gum arabic, is divided into six pills, one of which must be taken three times a day.

Besides the articles thus briefly examined, we might notice *salicine*, *lupuline*, *cornine*, *delphine*, *piperine*, *emetine*, and others; but we have not room for the examination, and, therefore, refer the reader to the books of *Materia Medica*.

**OF OILS AND RESINS.** These substances have some importance in a chemical point of view, and as all of them contain carbon, oxygen, and hydrogen, I do not know a more suitable place for them, than that which is here allotted.

The general characteristics of oils are, an unctuous feel, inflammability, and insolubility in water. They have been divided into fixed and volatile; the former being comparatively fixed in the fire, and therefore giving a permanently greasy stain to paper, while the stain of the latter is removed by heat.

Fixed oils are generally inodorous, nearly insipid, and lighter than water. They are generally of a yellow color, but this may be removed by the action of animal charcoal. Heated to redness in close vessels, they are converted into gas, which serves the purpose of illumination. Burnt in the open air, they are changed to water and carbonic acid; they contain, within themselves, the elements of both these results.

Although the oils do not dissolve in water, they may be permanently suspended by mucilage or sugar, in the form of *emulsion*. It is in this way that the castor oil mixture is prepared; the oil being first carefully blended with sugar and powdered gum arabic, is easily suspended in the requisite quantity of water.

The fixed oils are obtained, generally, from seeds, by pressure, with or without heat, while the volatile or essential oils are the product of distillation.

The volatile oils have a penetrating odor, and an acrid taste, both of which are often agreeable, when diluted. They are nearly insoluble in water, but quite soluble in alcohol. They burn like the fixed oils in open air, and yield the same products.

If some common almonds be steeped in hot water, then bruised in a mortar, and forcibly pressed in a strong linen bag, a clear and colorless oil will escape, which is usually called the *oil of almonds*. It is precisely in this way that all the fixed or expressed oils are prepared.

Chevreul has pointed out two principles in fatty matters, called *stearine*, (which is the solid portion,) and *elaine*, (the fluid part). If we expose a small quantity of a fixed oil to a freezing mixture, these parts will be separated, the thinner remaining above and the solid part settling to the bottom. The *stearine* may be obtained pure, by compressing fat between folds of bibulous paper, so as to remove the *elaine*.

The combustibility of oils, and their disposition to absorb oxygen from the air, may be shown by a simple experiment. Soak some cotton in linseed oil, and lay it aside in a safe place, where no harm can result, if combustion should occur. The oil gradually absorbs oxygen from the air, and as it is spread over a large surface, a considerable degree of heat is excited, and frequently the whole takes fire, in the course of one or two days.\*

If some sulphuric acid be mixed with half its bulk of a fixed oil, the mixture will soon be blackened, owing to the deposition of carbon; at the same time, fumes of sulphurous acid will be evolved.

Combustion may be excited in several oils, by pouring nitric acid on them. The reaction, in these cases, is violent, and much gaseous matter is disengaged. If an ounce or two of oil of turpentine be placed in a saucer, and the same quantity of strong nitric acid be dashed upon it suddenly, a violent combustion will ensue. This experiment is more certain if a little sulphuric acid be added to the nitric, and the mixture be heated. The vessel containing the acids should be fastened to one end of a stick, not less than six feet long, to avoid accident. Exp.

The only fixed oil that is soluble in alcohol, is the castor oil.

\* In this way many fires have occurred, which have been charged upon the incendiary.

If these be mixed in a bottle and shaken, they will form a transparent and colorless liquid. If a portion of this liquid be thrown into water, the latter will combine with the alcohol, and the oil will be set at liberty. The solubility of the essential oils in alcohol is well known; the article called *essence of peppermint*, is formed by dissolving oil of mint in spirit of wine.

We have already noticed the absorption of oxygen by linseed oil, but we have to say, further, that when that oil, in mass, is exposed to the open air, (and especially if heat be superadded in the process,) the oxygen absorbed acidifies the mucilaginous matter of the oil, and thus converts it into what is called *drying oil*, a term well known to painters.

The conversion of fixed oils into inflammable gas, may be shown on a small scale, by means of a retort and a good chafing dish, or a spirit lamp. The heat decomposes the oil, and recombines its elements, so as to form the combustible gas, which may be collected in proper vessels.

Soaps are formed by the union of oils with the several alkalis, to be spoken of hereafter. If one part of the liquid ammonia of the shops be mixed with eight parts of olive oil, a soap will be formed, which, in pharmacy, is called the *volatile or ammoniated liniment*.

If oil be boiled with a solution of potash or soda, soap will be formed; the former makes *soft* and the latter *hard* soap. Well conducted experiments have shown, that the alkali does not combine directly with the oil, but that the latter is decomposed, the stearine and elaine being changed into the margaric and oleic acids, which form soap by uniting with the alkali.

We can easily detach the alkali from soap by means of an acid. To a solution of soap in water, add a little sulphuric or muriatic acid, and the alkali will join the acid, leaving the oil free.

The alcoholic solution of soap is used as a test of hard waters, or waters holding sulphate or muriate of lime in solution. If this solution of soap be added to distilled or soft water, no change will appear; but if we add it to a hard water, a white, curdy precipitate, will be formed. The acid in combination with the lime in the water, will unite with the alkali of the soap, and form a soluble compound, while the lime will join the margaric and oleic acids, forming an imperfect, because insoluble soap, which gives the curdy appearance to the fluid.

Fixed oils perform an important service in many pharmaceutical operations, and in none, perhaps, more signally, than in the formation of *plasters*. Olive oil intimately combined with litharge, (an oxyde of lead,) constitutes the lead plaster of the shops,

which enters into the formation of several other plasters. Wax and resin, added in various proportions, give such a consistency to the mass as may be desirable.

Some of the oils have the faculty of dissolving phosphorus. Thus, if three ounces of olive oil and ten grains of phosphorus be mixed, and digested in a sand bath for a few hours, a part of the phosphorus will be dissolved. If the bottle containing this mixture be opened in a dark room, the whole will be luminous, owing to the sudden absorption of the oxygen of the air, by the phosphorus. Exp.

To show the rapidity with which oil of turpentine burns, let a small quantity be poured on a piece of paper, and lay upon it a grain or two of the salt called *chlorate of potash*, and touch this with a glass rod, dipped in sulphuric acid. The sulphuric acid decomposes the salt, combines with its base, and liberates an inflammable gas, (peroxyde of chlorine,) which instantly causes the oil of turpentine in the paper to burn. Common oil of turpentine contains but one part of oxygen in 100, and when it is made quite pure, no oxygen can be detected in it.

Closely allied to the essential oils, is the well known substance, *camphor*, furnished by the *laurus camphora*, and other plants. It differs from essential oils in composition, only in containing more carbon. Its specific gravity is less than that of water, and it is so volatile, that during warm weather, it evaporates very quickly, especially if the atmosphere be rather moist. If a lump, weighing several ounces, be placed in a wardrobe, in midsummer, the whole will disappear in a few weeks. Camphor is readily ignited, burns with a brilliant flame and much smoke; it melts at  $288^{\circ}$ , and boils at  $400^{\circ}$ .

To reduce camphor to the state of powder, a few drops of alcohol may be poured on it, and then the pestle will effectually triturate it. Alcohol is, of course, an active solvent of this substance, while water cannot dissolve more than a nine-hundredth part. If to an alcoholic solution, water be added, the solution is instantly destroyed; the water joins the alcohol, by virtue of superior affinity, and the camphor is reduced to its original state.\* Exp.

But alcohol is not necessary to the formation of camphor mixtures, that are to be administered internally. A good emulsion may be formed by grinding it well with three times its weight of almonds, and afterwards rubbing it with water, added in small quantities. Yolk of egg, the mucilages of starch and gum arabic, all answer the same end, and much better than sugar.

\* For this reason, camphor is best administered in the solid form. The moment the alcoholic solution enters the stomach, the camphor is precipitated.

Although simple water does not dissolve camphor, yet a surgeon at Madrid, discovered that highly carbonated water takes up a considerable portion.

Strong acetic acid dissolves camphor more freely than any other menstruum, but this solution could not be conveniently applied to internal use.

A very good plan is to rub the camphor in a mortar, with oil of olives, or oil of almonds, until the whole is closely blended; it can then be mixed, very readily, with mucilage of gum arabic, and afterwards, with such liquids as may be desirable. It has, likewise, the property of uniting with gum-resins, and of converting them into permanently soft and uniform masses, which may be incorporated with water.

An artificial camphor may be prepared, by passing a current of muriatic acid gas (to be noticed presently) through oil of turpentine; the compound thus formed, is of a crystalline texture, and has the odor of camphor, but may be easily distinguished from it, and, especially, by its insolubility in acetic acid.

RESINS differ from oils, in several respects. They are the inspissated juices of plants, and commonly occur either pure or in union with an essential oil. At ordinary temperatures, they are solid, brittle, inodorous, and insipid. They are, uniformly, non-conductors of electricity, and when rubbed together, become negatively electric. They are generally of a yellow color, and semi-transparent.

Resin contains the usual elements of vegetable matter, and, as analysed by Gay Lussac and Thenard, it yielded the following proportions, in 100 parts.

Carbon	75.944
Oxygen	13.337
Hydrogen	10.719
	<hr/>
	100.000
	<hr/>

From this composition, it will at once be obvious, that resinous matters are combustible. They burn freely in open air, giving a yellow flame with much smoke, and being resolved into carbonic acid and water. Heated in close vessels, they yield an empyreumatic oil, much carburetted hydrogen gas, with a small residue of charcoal.

Alcohol, ether, and essential oils dissolve resin, and it is precipitated from these solutions again, on the addition of water, because it is insoluble in the latter fluid. The best sol-



vent of resin is pure potash or soda, and it is also soluble in alkaline carbonates, by the aid of heat. In each case, the product is a saponaceous compound, which may be decomposed by acids.

To show the action of acids on resin, reduce some to a fine powder, and throw a spoonful of it into a deep glass containing an ounce of nitric acid. Copious ruddy vapors will be evolved, of a deeper color, and more dense, than are usually disengaged by the action of other inflammables on this acid. As pure resins are not soluble in water, but precipitated by it, from the alcoholic solutions, we see a reason why resinous tinctures should be made with pure spirit. The articles called *gum resins*, being compounded of gum and resin, require a mixed spirit for their solution, because the gum is perfectly soluble in water.

Such, then, is the general character of resins; and as their number is extensive, and some of them very important in pharmacy and the arts, the reader will do well to examine their history more in detail, in the larger works.

#### SECTION IX.—OF BI-SULPHURET OF CARBON.

This is also called carburet of sulphur, and has been but recently investigated. It is the last compound of carbon that we are to notice in this place.

The prefix *bi* denotes that there are two equivalents of sulphur, in union with one of carbon, making its entire equivalent, by weight, 38. Its specific gravity is 1.272, and its boiling point 110°, Fahrenheit.

To prepare this compound, the following directions are given. Take a porcelain or earthen tube, of about two feet in length, and an inch in diameter; coat it well with clay, and wrap it round with iron wire, to secure the coating in its place. The tube is next to be filled with fragments of charcoal, leaving sufficient room, however, for the free passage of vapor. We then pass the tube through a sheet iron or other furnace, so as to expose its central part to the action of the fire. A retort, one-third full of sulphur, is then fitted to one end of the tube, and well luted, taking care, at the same time, to support it, securely, by the retort stand. The lute, to be used, may be made of clay and sand, well mixed; and after applying this, it is to be covered with some fine plaster of Paris. To the other end of the tube, we next attach a bent glass tube, of half an inch in diameter, the outer end of which passes into one of the apertures of a

Woulfe's bottle, the latter being furnished with a safety tube, and being filled with water. It is necessary that these parts of the apparatus be also luted, carefully.

Every thing being thus duly adjusted, the fire is to be kindled in the furnace, and the porcelain tube brought gradually to a red heat. At this juncture, we apply an Argand's spirit lamp to the retort containing the sulphur, by which means we drive over the vapor of the sulphur and bring it into contact with the heated charcoal. Two equivalents of the vapor of sulphur combine with one equivalent of the carbon, forming the bi-sulphuret of carbon, which is condensed in drops that fall to the bottom of the Woulfe's bottle. The use of the safety tube is to prevent water from passing back into the heated tube, an accident which would impair the success of the operation. The charcoal used in this process, should be well prepared, and not mixed with any undecomposed, woody fibre.

The bi-sulphuret of carbon is never obtained quite pure, at first; but by adding a little of the article called chloride of calcium, in order to absorb any moisture that may be present, and then distilling, we render it sufficiently pure for all experimental purposes. The temperature to be observed in the process, must not exceed  $110^{\circ}$ , because beyond that heat, the whole would be changed to vapor. When pure, this article is in the form of a limpid and colorless liquid, remarkably transparent, having a very offensive and fetid smell, and an acrid, pungent taste. It is a very volatile substance, evaporating rapidly at natural temperatures, and producing a great degree of cold. It is highly inflammable, and burns with a bluish flame. Its vapor detonates violently with oxygen gas, and with the deutoxyde of nitrogen it burns very rapidly, but does not detonate, producing a very brilliant and dazzling light. These experiments are best performed, by filling a detonating bottle quite full of oxygen gas, then putting in a few drops of the bi-sulphuret and shaking the bottle, (first having corked it tightly,) till the whole is volatilised; the mixture is then inflamed, by means of a burning taper, or red hot wire.

Alcohol and ether combine with this substance, and it also dissolves sulphur, phosphorus, and iodine; but it is decomposed by chlorine. It unites with the alkalis, forming compounds which have been termed, *carbosulphurets*; and when shaken with an alcoholic solution of potash, a new acid is produced, called the *hydroxanthic*, from *xanthogen*, a word derived from the Greek, and signifying *yellow*. This appellation is based on the fact, that the compounds formed by this acid with several metals, are of a yellow color.

As yet, this article has not acquired much reputation in the practice of medicine. I find, however, a notice of its remedial powers, in the seventh volume of the *Journal de Pharmacie*. It is there especially recommended in the treatment of chronic gout and rheumatism. Dr Wutzer makes a tincture, by adding two drachms of the bi-sulphuret to a half ounce of alcohol; dose from five to ten drops, three times a day. A liniment is employed, at the same time, into which the bi-sulphuret also enters.

## CHAPTER VIII.

### Of Boron.

THIS is another of those bodies usually denominated simple, non-metallic. Its equivalent or combining number is 8.

Boron was prepared by Sir H. Davy, who discovered it, by submitting boracic acid, a compound of boron and oxygen, to the action of a galvanic battery. Gay Lussac and Thenard improved the process, and procured it in a greater quantity, by exposing boracic acid to a red heat, in a copper tube, with its own weight of potassium, the latter attracting oxygen from it and forming potash, which is easily removed by washing it with water; the boron remains in the solid form, of a dark olive color. The boracic acid ought previously to be deprived of water, as completely as possible, by protracted fusion in a platina crucible; but as it soon attracts water from the air, a detonation takes place at the instant of reduction, from the potassium reacting, at the same time, on the water combined with the acid and disengaging hydrogen gas. To avoid this, Berzelius recommends the dry borofluate of potash to be used instead of boracic acid; it is prepared by adding a solution of the fluuate of potash to a solution of borate of potash, and heating the gelatinous precipitate that is thrown down, till it assumes the form of a fine white powder. The theory of the action is the same as in the preceding process, the potassium taking the oxygen from the boracic acid in the compound salt.

Boron is insoluble in water and alcohol, undergoes no change when exposed to the air at ordinary temperatures, but inflames suddenly when heated to  $600^{\circ}$ , and is converted into boracic acid by combining with the oxygen of the air. In oxygen gas it burns more brilliantly, and attracts oxygen from a number of

substances that afford this element readily, as the sulphuric and nitric acids, the nitrate and chlorate of potash.

The only compound formed by boron with oxygen, is the *boracic acid*. Sixteen parts (two equivalents) of oxygen, and 8 (one equivalent) of boron, form one equivalent of this acid, or 24. This is called the *dry acid*, to distinguish it from that which is crystallised; the latter contains 24 parts of dry acid and 18 (two equivalents) of water. Its specific gravity is 1.479. It is soluble in water and in alcohol.

To prepare *boracic acid*, dissolve an ounce or two of crystallised borax (a compound of water, boracic acid, and soda) in four times its weight of boiling water, and add sulphuric acid previously diluted with four or five parts of water, till the solution becomes sensibly acid, (using a test-paper\* to ascertain when this is the case,) and then set it aside to crystallise. The sulphuric acid combines with the soda, forming sulphate of soda, which remains in solution, and crystals of boracic acid are deposited, which are purified by placing them on a paper filter, and washing them with cold water to remove any sulphate of soda that may be mixed with them. A minute portion of sulphuric acid is apt to adhere to them still, which can be removed only by repeated solution and crystallisation, or fusing it in a platina crucible; in all ordinary experiments, its presence is of no consequence.

Boracic acid crystallises in thin scales, which have a shining appearance. Its taste is sour and bitter; it reddens the vegetable blues, but produces the same effect on turmeric paper as the alkalis, turning it brown, as Mr Faraday pointed out. Crystallised boracic acid loses its water of crystallisation when exposed to a heat slowly increased, and the dry acid which remains is fused, and forms a transparent and colorless glass, on cooling. When a solution of boracic acid in water is boiled, or the crystallised boracic acid exposed suddenly to a high temperature, a considerable portion of it is carried along with the vapor of the water, but the dry acid may be exposed to a white heat, without being volatilised. From the facility with which it is fused, and the power which it has of communicating this property to its compounds, it is much employed, both in its pure state and in combination with soda, as a flux. The nature and use of fluxes, will be noticed when we treat of the alkalis.

\* For the mode of preparing test-papers, see the chapter on *coloring matters*.

## CHAPTER IX.

## Of Chlorine.

THIS term, derived from a Greek word signifying *green*, is applied to a gaseous fluid, of a greenish hue. It is here placed among the simple, non-metallic bodies, because it is now regarded as an undecomposed substance. Scheele, who discovered it, called it *dephlogisticated marine acid*; that is, marine or muriatic acid deprived of phlogiston. If by phlogiston, he meant hydrogen, as some modern chemists have contended, then it would seem that his views accorded precisely with those which now prevail, in relation to this interesting article. Berthollet and others called it *oxygenised muriatic* and *oxymuriatic acid*, on the supposition of its compound character; and these appellations prevailed until discovery prepared the way for the views which are now almost universal on this subject. Indeed, the late Mr Murray, even after Sir Humphrey Davy's announcement, contended most earnestly for the old theories, in respect of oxymuriatic acid; and Gay Lussac and Thenard, about the same time, came to the conclusion that it might be regarded, either as a simple or a compound. Indeed, our own views were far from being settled at the time of Mr Murray's controversy, and we were disposed to believe his statements in preference to those of Sir Humphrey. But the more carefully we have examined the subject, the more entirely have we been satisfied, that the simple or elementary constitution of chlorine is supported by the most decisive evidence, and that the strictest rules of philosophising force us to regard it as a simple body. In saying this, it must not be supposed, that we regard chlorine as absolutely simple. We believe that the time will come, when experiments, of the most conclusive sort, shall prove its compound nature; but at present, we must look at things as we find them, and patiently wait for the developement of that which is now hidden from us.

Notwithstanding all the light that has been cast upon this subject, it is a truth which, though strange, is undeniable, that every fact connected with the chemical history of this substance, is explicable on the old theory of its compound nature, as well as on the modern view of its simple character; and in some instances, the balance of evidence appears to be in favor of the old view of its constitution.

I have always objected to the appellation, *chlorine*, even admitting the simple character of the substance to be forever






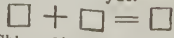
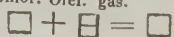
settled. I do not know, that a gas may ever be discovered which shall have precisely the same shade of green, but it is quite possible; and in the event of such a discovery, there would necessarily be some little confusion. For the present, however, the name now in use may answer all necessary purposes.

In whatever light we regard chlorine, it is one of the most important subjects in the whole range of chemical science.

Its combining or equivalent number, by weight, as deduced from analysis, is 36; by volume, it is one whole measure. Its specific gravity is 2.5. One hundred cubic inches weigh 76.25 grains. A pressure of four atmospheres changes it from a gas to a liquid; so also does the cold produced by the evaporation of sulphurous acid.

The following table, showing the equivalents of the most important compounds which chlorine forms with the non-metallic substances, will be found convenient to refer to, while studying its combinations. It will be seen that the equivalents, by weight as well as by volume, are here exhibited. Thus, the protoxyde of chlorine is a compound of chlorine and a non-metallic substance, viz. oxygen, in the proportion of 36 chlorine, and 8 oxygen, by weight; and a whole measure of the former, to a half measure of the latter, by volume.

*Compounds of Chlorine.*

	Chlor.	Oxyg.		Chlor.	Oxyg.	
Protoxyde of chlorine	36	+ 8 =	44	or		*
Peroxyde of chlorine	36	+ 32 =	68	or		
Chloric acid	36	+ 40 =	76			
Perchloric acid	36	+ 56 =	92			
	Chlor.	Hyd.		Chlor.	Hyd.	
Muriatic acid	36	+ 1 =	37	or		
	Chlor.	Nitrog.				
Chloride of nitrogen	144	+ 14 =	158			
	Chlor.	Sulph.				
Chloride of sulphur	36	+ 16 =	52			
	Chlor.	Carb.				
Chloride of carbon	36	+ 6 =	42			
Subchloride	36	+ 12 =	48			
Perchloride	108	+ 12 =	120			
	Chlor.	C. Oxyde.		Chlor.	C. Oxyde.	
Chlorocarbonic acid	36	+ 14 =	50	or		
	Chlor.	Olef. gas.		Chlor.	Olef. gas.	
Hydrocarbu. of chlor.	36	+ 14 =	50	or		
	Chlor.	Water.				
Hydrate of chlorine	36	+ 90 =	126			

\* One measure and a quarter.

The volumes represent the equivalents, by measure, of some of the compounds that have been examined in the gaseous form, and correspond with the equivalents, by weight, placed before them.

The best method of preparing chlorine, is by mixing one part of the peroxyde of manganese with four times its weight of muriatic acid, in a glass retort, collecting the gas in wide-mouthed bottles placed on the shelf of the pneumatic trough. The manganese should be reduced, previously, to a fine powder, and the water in the trough and in the bottles heated to the temperature of  $90^{\circ}$ , to prevent it from absorbing a large quantity of the chlorine. The retort should not be filled more than half full; the gas begins to be evolved whenever the materials are mixed, and on applying a gentle heat by a lamp or chauffer, it is disengaged more rapidly. Each bottle should have its stopple introduced under water whenever it is full, first drawing the finger round the edge with a little tallow, that it may be taken out easily afterwards. Common pneumatic jars may be employed to collect the gas when it is to be used immediately, but it must be recollected that it will be all absorbed, if it is left for a long time over water.

In this process, one equivalent of chlorine is obtained from every two of muriatic acid and one of peroxyde of manganese, the materials reacting on each other, in these proportions. Muriatic acid is composed of one equivalent of hydrogen and one of chlorine, and the peroxyde of manganese may be regarded as a compound of one equivalent of the protoxyde and one of oxygen. During the reaction which takes place, the protoxyde of manganese combines with one proportion of muriatic acid, forming muriate of manganese; the excess of oxygen at the same time combining with the hydrogen of the other equivalent of muriatic acid and disengaging the chlorine. The annexed diagram gives a more precise view of the combinations and decompositions which take place.

Before decomposition.		After decomposition.	
37 muriatic acid	{ chlorine hydrogen	36	36 chlorine.
37 muriatic acid		1	9 water.
44 perox. mang.	{ oxygen protox. mang.	37	73 mur. of mang.
		8	
		36	
<hr/> 118 <hr/>		<hr/> 118 <hr/>	<hr/> 118 <hr/>

As the common liquid muriatic acid contains more than its own weight of water, even in its most concentrated state, it is necessary to make allowance for this, and hence the large quantity which is directed to be mixed with the manganese, in the preceding paragraph. An ounce of the manganese, with the proper quantity of acid, will give a sufficient quantity of chlorine for the greater number of the experiments, usually performed with this substance; wide-mouthed bottles, capable of containing from 6 to 10 or 12 ounces of water, being employed to receive it.

The above diagram gives the rationale, according to the present doctrine, in relation to chlorine; but the process is equally explicable, on the ground of the old oxymuriatic theory. Thus, according to the Lavoisierian solution, the peroxyde of manganese is decomposed; its protoxyde combines with one portion of muriatic acid, forming muriate of manganese, while the oxygen arising from the decomposition of the peroxyde, combines with another portion of muriatic acid, and so gives rise to the oxygenised or oxymuriatic acid.

Another process for preparing chlorine consists in mixing 44 parts of the peroxyde of manganese, intimately, with 60 of chloride of sodium (dried common salt, a compound of chlorine and sodium) in a mortar, and pouring 98 parts of sulphuric acid on the mixture, previously diluted with half its weight of water and allowed to cool. Three or four hundred grains of salt will be a sufficient quantity, on the small scale, using a proportional quantity of the other materials, and placing them in a retort or flask with a bent tube adapted to it; the apparatus used for the preparation of hydrogen gas does very well, when a larger quantity is employed, supporting it on a retort stand, that heat may be easily applied; the sulphuric acid is diluted, to prevent the copious disengagement of muriatic acid fumes which always takes place, when strong sulphuric acid is poured upon the chloride of sodium.

It will be observed, that the materials are directed to be taken in the proportion of two equivalents of sulphuric acid, to one of the peroxyde of manganese, and one of the chloride of sodium. One equivalent of the acid, reacting on the peroxyde of manganese, forms sulphate of the protoxyde and disengages one equivalent of oxygen; this combines with the sodium of the chloride, forming soda, which immediately unites with the other equivalent of sulphuric acid, producing sulphate of soda, while the chlorine is disengaged. In the following diagram, representing the decomposition, the quantity of real sulphuric acid is stated, without the water that is usually combined with it, as it is not decomposed in the present instance.

Before decomposition.		After decomposition.	
60 chloride of sodium	{chlorine sodium	36 24	36 chlorine.
44 perox. mang.	{oxygen protox. mang.	8 36	
40 sulphuric acid		40	72 sulph. of soda.
40 sulphuric acid		40	76 sulph. mang.
<hr/> 184		<hr/> 184	<hr/> 184

Chlorine gas has a greenish-yellow color; a pungent, suffocating odor, even when diluted with a large quantity of air; and a disagreeable, astringent, metallic taste. In operating with it, care must be taken not to allow any to mix with the air, as it has often produced a degree of irritation in the lungs, and an anxiety and difficulty of breathing, with a total inability of taking a full inspiration, which has lasted for days, even when a single inspiration has been made, though still mixed with several times its bulk of air. In larger quantity, it produces a sense of strangulation, with a discharge from the nostrils; and in manufactories where large quantities are prepared, some of the workmen have occasionally fallen down quite senseless in an instant, when they have been exposed accidentally to a current of the gas. In these cases, they are removed immediately to the open air, and generally recover, very quickly, on dashing cold water upon them. It appears, too, that some of those who have been affected in this manner, never experienced any of the bad effects that accompany the inspiration of the gas, when diluted with air; probably from the great irritation it occasions in a pure state, causing a complete spasm of the glottis, and preventing any of it from passing into the lungs; the individual, therefore, suffers from the temporary suspension of respiration alone. Where the gas has got into the lungs, it has been recommended to take a drink, composed of the water of ammonia diluted with a very large quantity of water, to respire a little ammoniacal gas by keeping the mouth over some ammonia diluted with a less quantity of water, and to inhale the vapor of ether; if nothing else can be procured at the moment, considerable relief may be obtained, by holding the head over a large jar half full of hot water, and breathing into it.\*

\* The effect of habit in relation to this gas, is truly remarkable. Some persons become so much accustomed to it, as to care very little about it; while a stranger, on entering a manufactory containing it, would be almost suffocated in an instant.

A chemical manufacturer at Belfast has declared, that his workmen

Water absorbs one and a half times its volume of chlorine gas at the temperature of  $68^{\circ}$ , according to Thenard. The solution is called, in common language, *liquid* or *aqueous chlorine*; it has the color, taste, and smell of chlorine; it is stimulant and antiseptic, and gives chlorine gas when exposed to heat. This solution is most easily prepared in the manner directed for obtaining carbonic acid water; or the gas may be transmitted through water placed in the third bottle of Woulfe's apparatus, a small quantity of water being placed in the first and second to retain any muriatic acid that may be disengaged along with it. The *aqua oxymuriatica* of the colleges, is merely chlorine water.

Expose a portion of this liquid to the temperature of  $32^{\circ}$ , by placing it in ice-cold water, or surrounding it with a freezing mixture; it soon begins to congeal, and affords a solid mass of a yellow color, composed of chlorine and water; crystals of a definite compound of these substances, are obtained easily, free from any excess of water, by dropping a small quantity of this liquid into a bottle filled with chlorine gas, and placing it in a freezing mixture, or in the dark, at a temperature below  $32^{\circ}$ , for a few days. They form dendritical crystals on the side of the bottle.

Aqueous chlorine has no acid properties, but when exposed to the light, part of the water is slowly decomposed, one portion of the chlorine combining with the hydrogen and forming muriatic acid, while the other combines with the oxygen, and is converted into chloric acid. From the great attraction that subsists between chlorine and hydrogen, this liquid communicates oxygen to a number of metals and other substances which have an affinity for this element, the chlorine uniting with the hydrogen of a portion of water which is decomposed, and liberating oxygen.

Chlorine is particularly distinguished by its power of destroying all vegetable and animal coloring matters, and decomposing effluvia produced by contagious diseases, or arising from vegetable and animal matter in a state of putrefaction. It is accordingly employed in large quantity for bleaching and fumigation, and it is used for those purposes, either in the gaseous state, or in combination with lime or an alkali, and dissolved in water.

labor with impunity in an atmosphere of chlorine, where he could not remain above two or three minutes. The chief difficulties encountered, are acidity of stomach and similar complaints, which the men generally correct by taking chalk. One man has worked forty years in the establishment, and died finally in his eightieth year.



Aqueous chlorine is also frequently employed on the small scale, and a number of experiments may be made with the solution already directed to be prepared, by pouring it into solutions of vegetable coloring matter, as litmus, turmeric, indigo; and to others in a state of putrefaction, when their color and odor will be completely destroyed, if it be added in sufficient quantity. The method of making chloride of lime and Labarraque's disinfecting soda liquor, will be described under their respective bases.

Suspend some colored flowers in a bottle of chlorine gas, after putting in a few drops of water; in a short time they will, in general, have become of a pure, white color; but some are much more speedily deprived of color than others. Exp.

When chlorine is perfectly dry, vegetable coloring matter is not at all affected by it; and from a variety of experiments made with substances of this nature, it appears that chlorine acts principally by decomposing water, combining with its hydrogen and forming muriatic acid, while the oxygen that is eliminated, acts directly on the coloring matter and destroys it.

A number of inflammable substances burn in chlorine gas, and many of them take fire when mixed with it at natural temperatures.

Light a candle, suspended by a wire, and put it into a bottle of chlorine gas. It will continue to burn, but with a dull, red flame, and a large quantity of carbon will be deposited. Exp. The combustion is sustained, by the chlorine combining with the hydrogen of the inflammable matter, while the carbon is precipitated.

Pour some oil of turpentine on the lower part of a piece of thin, gray paper, folded into a match, allowing any excess to drop off, and then put it into a bottle of chlorine, holding it with a pair of pincers. The oil of turpentine will immediately take fire and burn with a lurid flame, the same reaction taking place as in the preceding instance, and with a similar deposition of carbon. Exp.

Mix half a measure, or one equivalent, of olefiant gas (a few cubic inches of the gas or a much larger quantity may be employed) with a whole measure or one equivalent of chlorine, in a glass jar, and apply a light to the mixture; it will burn quickly, with a flame similar to what is produced by the combustion of oil of turpentine in chlorine, the hydrogen combining with the chlorine and forming muriatic acid, while the carbon is precipitated. (See hydrocarburet of chlorine.) Exp.

Dry a piece of phosphorus, by means of cotton or blotting paper, and having placed it in a small copper cup, attached to

a wire, dip it into another bottle of chlorine. It takes fire instantly, and continues to burn for a considerable time, with a pale flame, combining with the chlorine and forming bi-chloride of phosphorus; a grain or two will be quite sufficient for a small bottle of the gas.

Exp.

All the metals can combine with chlorine; many of them take fire in this gas at natural temperatures, when introduced into it in a minute state of division, and the greater number of them when exposed to heat, with the exception of gold, silver, cobalt, and nickel.

Exp.

To see the phenomena attending the combination, put some leaves of Dutch gold (a compound of copper and zinc) into a copper cage, and introduce it into a bottle filled with chlorine; it immediately inflames, and chlorides of copper and zinc are formed.

Exp.

Throw some antimony or arsenic, reduced to a fine powder, into another bottle of chlorine; the metal immediately inflames and combines with the chlorine.

Exp.

Put some mercury into a copper cup, rubbed over with a little gas lute, to prevent the metals from combining, and place it in a bottle of chlorine, after heating it in the flame of a spirit lamp. The mercury takes fire, burning with a reddish-colored flame, and is converted into bi-chloride of mercury.

Chlorine can decompose many of the metallic oxydes (including the alkalis and earths) at a high temperature; and, in general, an equivalent of oxygen is disengaged for every equivalent of the oxyde that is decomposed, one equivalent of a metallic chloride being, at the same time, formed. The metallic oxyde to be decomposed, is placed in a coated porcelain tube, which is made to traverse a furnace in the usual manner, the chlorine being prepared in a flask retort, and passing over fragments of fused chloride of calcium, (to remove any water before it comes in contact with the oxyde,) placed in a small globe or tube, attached to its beak, and luted with plaster of Paris to the porcelain tube. The oxygen evolved is collected in a jar over the pneumatic trough, to which it is conducted by a bent glass tube fixed to the other extremity of the porcelain tube. The current of chlorine should be passed slowly but steadily over the oxyde, and it is better to employ a tubulated retort.

Chlorine is detected by a solution of the nitrate of silver, which gives a dense, white, curdy precipitate, composed of chlorine and metallic silver; this precipitate is soluble in ammonia, but insoluble in acids, and becomes of a dark color on exposure to the light.

The compounds of chlorine which are not acid, are termed *chlorides*, in England and this country, and *chlorurets*, in France.

Some attempts have been made, to introduce chlorine into medical practice, but it is not likely to be useful, excepting, merely, as an antidote. Its use as a tooth wash, and to correct fetid odors, to prevent or arrest the process of putrefaction, and to obviate the unpleasant smell consequent on animal decomposition, are all dependent on its chemical affinities. Thus, for example, it destroys the offensiveness of sulphuretted hydrogen, (the gas of privies,) by combining with the hydrogen of that compound, and precipitating the sulphur. This is precisely the rationale of the operation of the chloride of lime, now so largely employed to correct the stench of privies, sewers, drains, &c.

It is, no doubt, on the same decomposing power of chlorine, that it has succeeded in preventing the spread of epidemics. Thus we are told, that the workmen at a chemical manufactory at Belfast, from 1816 to 1819, were exempt from the fevers which desolated Ireland in that period, and that a large part of their operations is connected with chlorine. To the same agency of chlorine, must also be attributed its powers as an antidote to hydrocyanic (prussic) acid. It takes the hydrogen from that acid, and so destroys its poisonous character.

#### SECTION I.—PROTOXYDE OF CHLORINE.

Although there are four compounds of chlorine and oxygen gas, only two of them have acid properties, the remaining two being oxydes, in the strict sense of the term. The protoxyde is composed of an equivalent of chlorine, 36, and one of oxygen, 8, making its equivalent, by weight, 44; while, by volume, it is one and a quarter measure. Its specific gravity is 2.4. One hundred cubic inches weigh 74.5 grains.

The protoxyde of chlorine is prepared, by pouring one part of muriatic acid, diluted with an equal weight of water, on two parts of the chlorate of potash, (one or two drachms will be quite sufficient,) and collecting the gas that is disengaged, over the mercurial trough, applying a very gentle heat by a small spirit lamp. Great care ought to be taken in preparing this gas, as it explodes violently when exposed to a moderate heat, though nothing is mixed with it; the spirit lamp should be held immediately below the retort, so as not to play on its sides, and the gas should then come slowly away, producing a very moderate effervescence. A portion of chlorine is always disengaged along with it, and is removed by the mercury which immediately combines with it, but does not affect the protoxyde.

As this gas is much heavier than common air, it may be obtained, though not very pure, by displacement; that is, by passing it into a bottle, driving out the common air by its gravity, and finally filling the bottle. For this purpose an oil flask is furnished with a bent glass tube, one leg of which being in the flask, the other is fixed in the jar destined to receive the gas. When the materials, above named, have been placed in the flask and the gas begins to come over, it traverses the tube and soon occupies the lower part of the bottle. The deep color of the gas apprises us of the filling of the bottle or jar, which should be removed, and another put in its place, to be filled in like manner. In this process, the protoxyde of chlorine is formed, by the mutual action of the muriatic and chloric acids. One part of the muriatic acid employed combines with the potash of the chlorate of potash, forming muriate of potash, and at the same time liberating chloric acid, which is one of the compounds of oxygen and chlorine. Another portion of the muriatic acid (a compound of chlorine and hydrogen) reacts on the chloric acid in its nascent state; the hydrogen of the former combines with part of the oxygen of the latter, while its chlorine unites with another portion of oxygen, and so produces a part of the protoxyde which is evolved. The chloric acid, thus robbed of most of its oxygen, ceases to be an acid, and is reduced to the state of protoxyde of chlorine. The following diagram is intended to explain this process. It represents the mutual action of muriatic acid and chloric acid, and supposes the latter to have been liberated from the chlorate of potash, by the action of a portion of muriatic acid.

Before decomposition.		After decomposition.	
2 equivalents of muriatic acid equal to 74	hydrogen 1	9 water.	
	hydrogen 1		
	chlorine 36		
	chlorine 36		
1 equivalent of chloric acid equal to 76	chlorine 36		
	oxygen 8		
	oxygen 8		
	oxygen 8	44	protox. of chlorine.
	oxygen 8	44	do. do.
	oxygen 8	44	do. do.
<hr/> 150 <hr/>	<hr/> 150 <hr/>	<hr/> 150 <hr/>	

There is always some free chlorine mingled with the protoxyde, which must proceed from the reaction of the muriatic and



chloric acids on each other, in different proportions. If a large excess of strong muriatic acid be mixed with the chlorate of potash, the chloric acid will lose all its oxygen, and nothing but chlorine will be evolved. Thus, if five parts of pure muriatic acid be added to one part of chlorate of potash, it follows that only one equivalent of chloric acid can be liberated. Now this one equivalent of chloric acid contains five equivalents of oxygen to one of chlorine, while the five parts of muriatic acid contain five equivalents of chlorine and five of hydrogen. The hydrogen of the muriatic acid and the oxygen of the chloric acid combine to form water, and nothing but chlorine remains.

The protoxyde of chlorine has a rich, greenish-yellow color, from which Sir H. Davy gave it the name of *euchlorine*; its odor resembles that of burned sugar. It has no acid properties, destroys the vegetable colors, and is absorbed in considerable quantity by water, which can take up about ten times its volume of this gas.

Introduce a red hot iron wire, bent at one end, or a lighted candle, into a strong tube or detonating bottle, filled with the protoxyde of chlorine; an explosion immediately takes <sup>Exp.</sup> place, and a flash of light is at the same time perceived. The detonation arises solely from the separation of the elements of this compound, in which they exist in a condensed state; and, certainly, no one could have anticipated, that this would have been accompanied by an evolution both of heat and light. If the explosion has been occasioned by a combustible substance, it will continue to burn afterwards, in the mixed gases. Every measure and a quarter of the gas expands, during decomposition, to a measure and a half, one measure of which (one equivalent) is chlorine, and the remaining half measure corresponds with the equivalent of oxygen.

Phosphorus takes fire in this gas, and with hydrogen a detonating mixture is formed. The proper proportions are two measures (two equivalents) of hydrogen to every measure and a quarter of the protoxyde; the mixture must be inflamed by an electric spark, or a lighted match; the products are water and muriatic acid.

## SECTION II.—PEROXYDE OF CHLORINE.

As there are but two oxydes of chlorine, we do not designate this as the deutoxyde, as we should do, if there were three. Being the highest or last state of oxydation, we use the prefix that is common in such cases, and therefore call it *peroxyde*.



The quantity of chlorine is the same as in the protoxyde, but the proportion of oxygen is quadrupled. Hence the peroxyde consists of 36 chlorine, (one equivalent,) and 32 oxygen, (four equivalents,) by weight; by volume, its equivalent is two whole measures. Its specific gravity is 2.36; and 100 cubic inches weigh 72 grains.

The peroxyde of chlorine, like the protoxyde, requires much care in its preparation, and it should be made only in small quantities, at a time. Its general properties are similar to those of the protoxyde, but it has a deeper color, explodes with greater violence, and at a lower temperature. Every two measures (one equivalent) expand to three measures, one of which is chlorine, (one equivalent,) and the other two consist of oxygen, (four equivalents).

It is obtained, by mixing powdered chlorate of potash (60 grains may be taken) with no more sulphuric acid than is neces-

Exp. sary to convert it into a solid paste; the mixture is put into a small tubulated retort, the body of which is placed in hot water, taking care not to allow it to arrive at a boiling temperature; it may be collected, as it is disengaged, over mercury, which has no action on it.

Instead of preparing a considerable quantity of this gas, in a retort, a portion sufficiently large to show its explosive property, may be made in the following manner. Take a glass tube, three or four inches long and an inch in diameter, open at one end. In the bottom, place one or two grains of the chlorate of potash,

Exp. and then add a few drops of sulphuric acid, and place the tube in a well-heated sand bath. This mode of heating answers better, in the present instance, than a spirit lamp, which is recommended by Mr Faraday, because it is always safe. In a few moments, slight cracks will be heard, and if a red hot iron wire be now introduced, the explosions will be violent. Sometimes the tube is blown to pieces, and, therefore, the iron wire should be fastened to one end of a long stick, and the hand, most exposed, should be covered with a glove.

To understand the theory of the process, it must be borne in mind, that chloric acid is a compound of five equivalents of oxygen to one of chlorine, or it may be regarded, as a compound of one equivalent of peroxyde of chlorine and one of oxygen. Two equivalents of sulphuric acid decomposing two of the chlorate of potash, two of sulphate of potash are formed, and two of chloric acid are disengaged. These react on another equivalent of chloric acid in union with potash; each of the two first equivalents of acid loses one equivalent of oxygen, and is thus changed into peroxyde of chlorine, while the two equivalents of

oxygen, thus detached, unite with the undecomposed equivalent of chloric acid, and form perchloric or oxychloric acid, which seizes the potash, and gives rise to the perchlorate or oxychlorate of potash. The following diagram will illustrate this complicated action.

Before decomposition.		After decomposition.	
2 equiv. of chloric acid	{ perox. chl. 68	68	perox. of chlorine.
	{ perox. chl. 68	68	perox. of chlorine.
	{ oxygen 8	92	perchloric acid.
	{ oxygen 8		
Chloric acid	76		
	<hr/> 228	<hr/> 228	
	<hr/>	<hr/>	

Peroxyde of chlorine has an aromatic smell, and none of the peculiar odor of chlorine. It is not affected, at common temperatures, by any of the simple inflammables except phosphorus, which inflames and decomposes it, burning brilliantly, afterwards, in the mixture of chlorine and oxygen that remains.

It is even capable of inflaming phosphorus under water, which may be easily shown by putting 30 or 40 grains of the chlorate of potash into a long, narrow glass, filling it with cold water and pouring sulphuric acid on the chlorate, through a funnel that reaches to the bottom of the glass, after throwing in a few grains of phosphorus, cut into thin slices; small quantities of acid must be added, at a time, and the peroxyde is disengaged, inflaming the phosphorus, as it rises through the water. This is a very pleasing experiment and easily performed. Exp.

### SECTION III.—CHLORIC ACID.

On the supposition, that the theory of chlorine is well founded, the term chloric acid must be correct, and indicates such a degree of oxygenation of chlorine, as converts it into an acid. It is the same acid that was formerly called *hyperoxymuriatic*, a name that was based on the compound theory of chlorine.

Chloric acid is prepared, usually, by decomposing the salt called chlorate of barytes, by means of sulphuric acid. The acid, being first diluted with ten times its weight of water, is added to a weak solution of the salt, as long as any precipitate is formed, care being taken not to add an excess of acid. The precipitate that falls, is sulphate of barytes, and the chloric acid remains in solution.

The equivalent number for chloric acid is 76, being composed of 40 (five equivalents) oxygen and 36 (one equivalent) chlorine.

Chloric acid reddens the vegetable blues, is decomposed by a number of substances which have a great affinity for oxygen, and forms an important class of salts, all of which are decomposed by heat, and distinguished by the violent action that takes place, when they are exposed to heat, along with inflammable matter; with many of them, it detonates at natural temperatures, by friction or percussion.

*Perchloric acid* is formed by the reaction of sulphuric acid on the chlorate of potash. It is also called *oxychloric acid*, and is estimated to consist of chlorine 36, (or one equivalent,) and oxygen 56, (or 7 equivalents,) making its equivalent number 92. Serullas formed it, by boiling concentrated chloric acid; in which operation, there was a disengagement of chlorine as well as oxygen, and an accumulation of oxygen in part of the chloric acid, that was not decomposed. He employs it as a reagent, to distinguish between the salts of potash and soda.

#### SECTION IV.—MURIATIC OR HYDROCHLORIC ACID.

This acid was for many years an opprobrium to our science. Every one was prepared with a conjecture respecting its true nature, but all were ignorant; and until the discoveries in relation to chlorine, we knew nothing concerning its real character. It was discovered by Dr Priestley, in 1772.

It was formerly called *acid of sea salt*, *marine acid*, *spirit of salt*, and the like; but no name seems to be so descriptive of its composition as that which the French employ, viz. *hydrochloric*. Its equivalent is 37, by weight, being composed of 36 chlorine and one hydrogen. By volume, its equivalent is two whole measures. Its specific gravity is 1.28; and one hundred cubic inches weigh 39.18 grains. Its natural state is gaseous. There are various directions for preparing liquid muriatic acid, but the following is, probably, as good a mode as can be pursued.

Common salt is exposed, for an hour or two, to a red heat in an earthen vessel, to decompose any nitrates which it may contain, (a small quantity of these salts being occasionally found in it,) and one part of it is mixed with an equal weight of sulphuric acid, previously diluted with a third of its weight of water, in a glass retort, pouring the acid upon the salt through a long funnel. The retort is then placed in a sand bath, and a receiver adapted to it, containing water equal in weight to two-thirds of

the salt employed. The materials should not occupy more than a third of the body of the retort, as they are apt to boil over, when exposed to heat. The furnace is kindled after every thing has been properly adjusted, and the distillation continued to dryness; the receiver is kept cold and the apparatus arranged in the usual manner, and when the atmospheric air in the retort and the receiver has been expelled, which usually takes place a short time after the mixture begins to boil, the receiver may be luted to the neck of the retort with a little clay, taking care always to keep it sufficiently cold, by a constant stream of water. One equivalent of sulphuric acid, 49, is capable of decomposing one equivalent of chloride of sodium,\* and giving an equivalent of muriatic acid; but Dr Hope found that when equal weights are employed, the decomposition, as it is usually effected, is more complete, and a larger quantity of muriatic acid more easily procured. The acid gas produced by the action, and the water previously mixed with the materials, are condensed in the receiver, and considerable heat is produced, not only by the condensation of the watery vapor, but also by the combination of the muriatic acid gas with the water in the receiver. The specific gravity of the liquid obtained is 1.117, and contains about 34 per cent. of dry acid.

In this process, each equivalent of muriatic acid is formed by one equivalent of the chlorine (of the chloride) combining with the hydrogen of the water in the common sulphuric acid, the sodium taking the oxygen and forming soda, with which the sulphuric acid at the same time unites, and is converted into sulphate of soda; the excess of sulphuric acid (which is not represented in the following diagram) combines with part of the sulphate of soda and forms a bi-sulphate.

Before decomposition.			After decomposition.	
49 common sulph. acid	{	hydrogen	1	37 muriatic acid.
		oxygen	8	
		dry acid	40	72 sulphate of soda.
60 chloride of sodium	{	chlorine	36	
		sodium	24	
<hr/> 109 <hr/>			<hr/> 109 <hr/>	<hr/> 109 <hr/>

Dilute some of the acid with an equal bulk of water, and drop into it a solution of the muriate of barytes; if any precipitation takes place, it contains sulphuric acid, from which it may be separated by a second distillation with a small quantity of the

\* A name for common salt, of which further notice will be taken in due season.



chloride of sodium. Bottles may be filled, without a mercurial trough, by displacement, in the manner described when treating of protoxyde of chlorine.

Muriatic acid may be formed, also, by detonating a mixture of equal measures of chlorine and hydrogen; the bottle should be filled half full of hydrogen first, and then filled up with chlorine, corking it immediately to prevent the absorption of any of the chlorine by the water. A lighted match is then applied in the usual manner, the mixture detonates with flame and a loud report, but no condensation attends the combination, two measures of muriatic acid gas being formed. Pour an infusion

Exp. of litmus into the bottle immediately after the detonation; it will be reddened by the muriatic acid; if it had been poured in before the combination was effected, the chlorine would have rendered it colorless.

Fill the detonating bottle again, with hydrogen and chlorine, in the same manner and in the same proportions, cork it, and expose it to the direct rays of the sun. A detonation takes place, and if the cork is not forced out, the muriatic acid gas will be completely absorbed, on taking it out under water. If the mixture is kept in the dark, no action takes place; and in the shade, the chlorine and hydrogen combine slowly, without detonation.

When muriatic acid is required in the gaseous form, it must be collected in jars or bottles over the mercurial trough, as it is instantly absorbed in large quantity by water. The easiest method of obtaining it, is by exposing a strong solution of muriatic acid in water (common liquid muriatic acid) to a gentle heat in a tubulated retort, capable of containing three or four ounces; it may be filled half full, and a spirit lamp, held in the hand, will be found the most convenient method of applying the heat. Muriatic acid gas is soon disengaged in large quantity, and may be collected after the air has been expelled from the retort; no water is distilled over along with it, till a considerable portion of gas has been expelled. A sufficient quantity should be collected, at once, for all the experiments intended to be performed with it.

Muriatic acid gas has an acrid, pungent, and suffocating odor, and a strong, acid taste, even when combined with a large quantity of water. It is transparent and colorless, produces fumes when mixed with the air, by combining with the watery vapor which it contains, and cannot support combustion or respiration.

Muriatic acid gas has a very great affinity for water, which can absorb 480 times its bulk of this gas, considerable heat being produced by the combination, and the specific gravity of the resulting liquid being 1.21.



Take a long tube filled with acid gas at the mercurial trough, close it with the thumb or finger, transfer it to a bason of water colored blue by an infusion of cabbage or litmus, and remove the finger under the surface of the water; the gas is immediately condensed, the colored water forced up into the tube with explosive violence by the pressure of the atmosphere, and reddened at the same time by the acid; if any common air has been mixed with the gas, the absorption goes on more slowly and the air remains in the tube. Exp.

Fill a small test-tube with water, and introduce a little into a jar of muriatic acid gas over the mercurial trough; observe the large quantity of gas it condenses, as indicated by the rising of the mercury. Exp.

Put a piece of ice into another jar; it is melted almost immediately, and a solution of muriatic acid formed. Exp.

Take 300 grains (or any other quantity) of the acid, procured by distillation in the manner described, dilute it with an equal quantity of water, and drop into it fragments of marble from a given weight of this substance, till it will not dissolve any more. Then ascertain the quantity dissolved, by weighing what remains, (washing with water and drying what may still be mixed with the liquid,) and calculate the quantity of real muriatic acid which it contains, allowing 37 per cent. for every 50 grains of marble dissolved; for 37 parts (one equivalent) of real muriatic acid act on 50 parts of marble, (a compound of carbonic acid 22+28 lime,) combining with the lime and disengaging the carbonic acid.

Fill a long tube half full of strong liquid muriatic acid, and pour water gently over it till the tube is full, then close the mouth with the finger, rubbed over with a little wax lute, and invert it till the two fluids are completely mixed. Heat is evolved, and when the mixture is cooled, it will be found to occupy a smaller volume than the liquids before they were combined. Exp.

Strong liquid muriatic acid is transparent and colorless, when perfectly pure, emits copious fumes on exposure to the air, boils at  $110^{\circ}$ , giving off muriatic acid gas, and freezes when exposed to a very low temperature.

It has usually a light, greenish-yellow color, from the presence of a small quantity of chlorine or oxyde of iron; the former is produced by a very minute portion of nitric acid (disengaged from some nitrates mixed with the salt by the sulphuric acid used in its preparation) reacting on part of the muriatic acid, and the latter also from some impurities in the materials employed, or from the iron vessels in which it is prepared on the large scale. Ammonia, added in excess, pre-

precipitates any oxyde of iron that may be present; chlorine may be occasionally detected by the smell, or by gold leaf, which it speedily dissolves, pure muriatic acid having no action on it.

The most delicate test of muriatic acid, is a solution of the nitrate of silver, which gives a copious, curdy precipitate of the chloride of silver, the hydrogen of the acid combining at the same time with the oxygen of the oxyde; chlorine, indeed, gives a similar precipitate, but its action on the vegetable coloring matter distinguishes it sufficiently from the acid.

Muriatic acid combines with the salifiable bases, and forms an important class of salts, termed *muriates*. They are, in general, decomposed by heat, the hydrogen of the acid uniting with the oxygen of the oxyde, the resulting products being water and a metallic chloride; they are decomposed, also, by sulphuric acid.

Muriatic acid has been much used in medical practice, as a caustic, astringent, and tonic. It has also been recommended, in form of gas, for the purification of infected places; but it does not seem calculated to answer this end, so well as chlorine, which is one of its component parts.

Mixed with nitric acid, in the proportion of two parts muriatic to one of nitric acid, it constitutes the *nitromuriatic acid* or *aqua regia*, which is the proper solvent of gold, and has also been employed for the purpose of a bath. In this mixture, there is a decomposition of both acids, with evolution of chlorine, which being held in solution, is the efficient agent in making the solution of gold, and, also, in accomplishing any favorable results ordinarily had from the acid bath.

The nature of the reaction effected by mixing these acids, may be shown by adding a half ounce of muriatic acid to one ounce of nitric acid, and exposing the mixture to  
 Exp. a gentle heat in a flask. Dense, red fumes will appear, and chlorine being evolved, will be absorbed by the water in the mixture.

That there is really no such thing as a chemical compound of nitric and muriatic acids, as such, may be demonstrated by a simple diagram; premising, that nitric acid may be regarded as a compound of nitrous acid 46, and oxygen 8.

Before decomposition.		After decomposition.	
37 or 1 equiv.	{ chlorine	36	36 chlorine.
muriatic acid		1	
54 or 1 equiv.	{ hydrogen	8	9 water.
nitric acid		46	
	{ oxygen		
	{ nitrous acid		46 nitrous acid.
—		—	—
91		91	91
—		—	—

When an excessive dose of muriatic acid has been swallowed, the same antidotes are proper, that were noticed when speaking of the poison of nitric acid.

As ammoniacal gas instantly combines with muriatic acid gas, forming dense, white clouds, either of those substances may be employed to detect the escape of the other, from vessels in which they are prepared.

We detect sulphuric acid, when present in muriatic acid, by adding a few drops of muriate of barytes, which occasion a white precipitate of sulphate of barytes. Iron is indicated, by saturating a portion of the acid with pure carbonate of soda, by which a carbonate of iron will be formed. If we then add a little of the test called ferrocyanate of potash, a blue precipitate will be formed, which is Prussian blue. Copper is detected by the addition of liquid ammonia, until the acid is entirely saturated; it produces a deep blue color.

We are not always able to decide on the nature of the impurity present, simply by the tinge of the acid. This may depend on vegetable extractive matter, on chlorine, or on iron.

#### SECTION V.—CHLORIDE OF NITROGEN.

This term is analogous to the oxydes of nitrogen, already examined; and not only is it of modern origin, but the substance designated by it, was unknown until after the simple nature of chlorine was announced. Its equivalent is 158, being composed of 144 chlorine and 14 nitrogen.

This compound detonates with great violence, when touched with many inflammable substances, or exposed to heat, and probably, also, from several other causes which have not yet been discovered, as many accidents have taken place during its preparation, which have been attended with very serious consequences.

It is prepared, by inverting a jar or wide-mouthed bottle (capable of containing about 12 or 14 ounces) full of chlorine, over a dilute solution of the muriate of ammonia, made by dissolving an ounce of the salt in ten or twelve ounces of water. One portion of the chlorine takes the hydrogen of the ammonia, forming muriatic acid; and the other, combining with the nitrogen, is converted into the chloride, which collects in the form of an oil, on the surface of the liquid, and drops into a very strong, shallow, leaden cup, on which the bottle is placed, resting on a plate containing the solution, previously heated to the temperature of  $90^{\circ}$ ; an additional quantity must be ready

to fill up the plate, as the absorption of the chlorine proceeds. Great care must be taken not to shake the bottle, and any gas lute or fatty matter adhering to it, must be removed by washing it with a dilute solution of potash, before it is filled with chlorine. When it has fallen into the leaden cup, the bottle is carefully moved from the cup to the plate, and the matter taken cautiously away.

The liquid remaining above the chloride in the cup, is withdrawn by dipping small pieces of filtering paper into it; and on touching the chloride with a drop of olive oil at the end of a stick at least two or three feet long, a loud explosion takes place, though the quantity may not exceed the bulk of a pea. It is resolved, at the same time, into chlorine and nitrogen; the manner in which inflammable substances act, when they cause an explosion with it, has not been ascertained. It is an exceedingly dangerous article to prepare, or to experiment with; and as mere gratification of curiosity, without any profit, is not sufficient compensation for personal injury, I would have all young operators to forego the hazardous pleasure connected with it, entirely.

Its odor is extremely penetrating and almost insupportable, affecting the eyes very much on leaning over it, even for a second or two; at natural temperatures it volatilises rapidly, and explodes when heated to  $200^{\circ}$ .

#### SECTION VI.—CHLORIDES OF SULPHUR, PHOSPHORUS, CARBON, BORON, HYDROCARBURET OF CHLORINE, CHLOROCARBONIC AND CHLOROCYANIC ACIDS.

*Chloride of sulphur* is prepared by transmitting chlorine over flowers of sulphur, in a coated glass or porcelain tube, moderately heated. It is a reddish-colored liquid, easily volatilised, and emits very acrid, irritating fumes. It decomposes water, alcohol, and ether, a portion of sulphur being deposited, and the remaining sulphur converted into sulphurous and sulphuric acids, by the oxygen which it takes from these substances, the chlorine, at the same time, uniting with the hydrogen and forming muriatic acid.

There are two compounds of chlorine and phosphorus, viz. the *chloride* and *bi-chloride*. The latter is formed, when phosphorus is introduced into chlorine, perfectly dry; the phosphorus inflames, and a white matter collects on the inside of the vessel, which is the per or bi-chloride. It is a compound of two equivalents of chlorine, 72, and one equivalent of phos-



phorus, 12. The bi-chloride of phosphorus and water mutually decompose each other, the two proportions of chlorine taking two of hydrogen from the water, and forming two of muriatic acid, while the phosphorus combines with two of oxygen, and is changed into phosphoric acid.

The *chloride*, or protochloryde may be formed by heating the bi-chloride with phosphorus, or by passing the vapor of phosphorus over corrosive sublimate, contained in a glass tube. It is a clear liquid, like water, emits acrid fumes when exposed to the air, owing to decomposition of watery vapor. On mixing it with water, mutual decomposition ensues, heat is evolved, and a solution of muriatic and phosphorous acids is obtained. It is a compound of one equivalent of chlorine, 36, and one of phosphorus, 12.

*Hydrocarburet of chlorine* is prepared by mixing two measures of chlorine with one of olefiant gas, and leaving the mixture over water. The gases combine slowly together, forming a few drops of an oily liquid, which must be washed with water, and purified by distillation, after mixing it with some chloride of calcium, to retain any water that may be adhering to it. Its taste is sweet; it boils at about  $150^{\circ}$ , and is completely decomposed by a red heat. It is also called *chloric ether*.\*

By subjecting this compound, repeatedly, to the action of chlorine gas exposed to the rays of the sun, Mr Faraday succeeded in obtaining a compound of chlorine and carbon, the hydrogen being removed and converted into muriatic acid.

It has received the name of *perchloride of carbon*; for the details of the process for preparing this compound, I must refer to Mr Faraday's paper in the Philosophical Transactions, for 1821. When its vapor is passed through a red hot porcelain tube, a large quantity of chlorine is disengaged, and a liquid is obtained which is composed of one equivalent of carbon and one of chlorine, the *protochloride of carbon*. Another compound of chlorine and carbon, the *subchloride of carbon*, has been described in the first volume of the new series of the Annals of Philosophy; it contains two equivalents of carbon, 12, and one of chlorine, 36.

*Chlorocarbonic acid* is a compound of chlorine and carbonic oxyde gas, discovered by Dr John Davy; it has been called *phosgene gas*, from its being formed by the action of light on the mixed gases. It is transparent and colorless, reddens litmus

\* Mr S. Guthrie, of Sacket's Harbor, manufactures chloric ether, in large quantities, by distilling a mixture of three pounds of chloride of lime and two gallons of strong alcohol. This mixture yields one gallon of the ethereal spirit. It has been introduced into medical practice. See Silliman's Journal, October, 1831.



paper, combines with ammonia, and is decomposed by water, when carbonic and muriatic acids are produced, the hydrogen of a portion of water combining with the chlorine, and the oxygen with the carbonic oxide.

*Chlorocyanic acid* is called, in some of the books, cyanuret of chlorine, and oxyprussic acid. The latter name was given by Berthollet, who first noticed the article, and supposed it to contain oxygen and prussic acid. Gay Lussac called it chlorocyanic acid, and procured it, by passing chlorine gas into an aqueous solution of hydrocyanic acid, until the liquid acquired bleaching powers; the excess of chlorine was then removed, by agitation with mercury, and then heating the mixture so as to expel the gaseous cyanuret of chlorine or chlorocyanic acid. The chemical changes that ensue are a little complicated; at first, the elements of hydrocyanic acid unite with separate portions of chlorine, and give rise to muriatic acid and chlorocyanic acid; and when heat is applied, the elements of the chlorocyanic acid and water react on each other, in consequence of which, muriatic acid, ammonia, and carbonic acid are generated.

Chlorocyanic acid is a compound of one equivalent of chlorine, 36, and one of cyanogen, 26. It has an offensive odor, irritates the eyes, corrodes the skin, and is highly injurious to animal life.

Chlorine is also susceptible of union with boron, and the combination is *chloride of boron*.

We have already noticed the production of boron, by Sir Humphrey Davy, and adverted to its combustibility. It takes fire in an atmosphere of chlorine, and emits a vivid light. Berzelius made use of the boron, previously heated, whereby it was rendered more compact, so that heat was required to excite combustion. When thus burnt in dry chlorine, in a glass tube, a colorless gas was formed, which was collected over mercury. This metal absorbed the free chlorine that was present, and left pure chloride of boron. It would not answer to make the collection over water, instead of mercury, because that fluid, as well as the gaseous chloride of boron would be decomposed; the hydrogen of the water and the chlorine of the chloride uniting to produce muriatic acid, while the boron and oxygen of the water would form boracic acid.

The composition of this chloride has been determined by its action on water, and is said to be two equivalents, 72, of chlorine, and one, 8, of boron; it is, consequently, a bi-chloride. Every equivalent of the compound and two of water, yield one equivalent of boracic acid and two of muriatic acid.

In addition to the process above named, for making chloride

of boron, it may be generated by the action of dry chlorine on a mixture of charcoal and boracic acid, heated to redness in a porcelain tube. The charcoal, in this case, abstracts all the oxygen from the boracic acid, and thus reduces it to the state of boron, which then unites with the chlorine, forming chloride of boron.

---

## CHAPTER X.

Of Iodine. \*

THIS is another of the substances usually regarded as simple, although its general properties and appearance are calculated to make us a little sceptical on this subject. It was discovered by M. Courtois, in 1812, while engaged in preparing the carbonate of soda, from sea-weeds. He observed that the metallic vessels were very much corroded by the residual liquor, and on adding a little sulphuric acid, he saw a dark colored matter thrown down, which was changed into a beautiful, violet vapor by the action of heat. He was not aware of the nature of this substance and gave a portion of it to Mr Clement, who described some of its leading properties in the Royal Institute of France, in 1813. But its true nature was not understood, until Gay Lussac and Davy ascertained that it was a simple, non-metallic substance, very analogous in its properties and habitudes to chlorine.

Its equivalent, by weight, is 124, and by volume, one whole measure. The specific gravity of solid iodine is 4.948; that of its vapor is 8.61. One hundred cubic inches weigh 262 grains. It is very insoluble in water, but dissolves readily in alcohol and ether.

The process, usually followed for the preparation of iodine, was proposed by Dr Ure. The liquid that remains after most of the saline matter has been extracted from kelp,† is heated to

\* This name, like chlorine, is given simply to denote the color of the substance, especially in its gaseous form. Its derivative expresses the *violet* color, so peculiar to *iodine*.

† Kelp is the term applied to the ashes of incinerated sea-weed, prepared on the coast of Scotland, and elsewhere, and is used in large quantities by many of the soap manufacturers, from whom the above liquid may be procured.

the temperature of  $230^{\circ}$ , poured into a stone-ware bason, (which should not be filled more than half full,) and an ounce, by weight, of sulphuric acid, previously diluted with an equal bulk of water, mixed with every eight ounces, by measure, of the liquid employed; a brisk effervescence immediately takes place, a large quantity of sulphur is at the same time precipitated, and crystals of the sulphate of soda are deposited when the mixture is cold.

The salts which this fluid contains, are composed principally of soda combined with sulphuretted hydrogen, carbonic, sulphurous and hydriodic acids, all of which are separated by the sulphuric acid, which combines with the soda and forms the sulphate, that is afterwards crystallised. The carbonic acid mixed with part of the sulphuretted hydrogen and sulphurous acids, produces the effervescence, as they are disengaged; the rest of the latter acids mutually decomposing each other, the hydrogen of the one combining with the oxygen of the other, while the sulphur of both is precipitated; the hydriodic acid which remains in the liquid, is separated from the sulphur by filtration through paper.

In the last stage of the process, the iodine is obtained by mixing 1000 grains of the peroxyde of manganese with every 12 ounces, by measure, of the filtered liquid, in a glass retort, (capable of containing at least 24 ounces when quite full,) condensing it in a receiver which is kept cold, in the usual manner. Dr Ure recommends a glass flask to be used, and the iodine, which is deposited in small crystals, to be condensed in a large globe or receiver placed above it.

This may be managed conveniently, by using a retort stand with two rings. On the lower ring, let the flask be placed, and have a round piece of tin, 10 inches in diameter with an inch hole in the centre, fixed on its neck. The glass globe or receiver must have a neck large enough to receive the neck of the flask, and when inverted, is to be kept in its place by the upper ring of the retort stand. The design of the tin plate is to prevent the receiver from being too much heated, and so soon as it becomes warm from the condensation of iodine and watery vapor, it is to be removed and another put in its place.

When the prepared iodine liquor has been kept some weeks before it is used, sometimes little or no iodine is obtained on heating it with peroxyde of manganese; in such cases, the addition of a little caustic potash, causes the iodine vapors to appear in as large a quantity, as if the liquid had been newly prepared. From 80 to 100 grains of pure iodine are obtained from this quantity of liquid.

The theory of the process is similar to the theory of the prep-

aration of chlorine from muriatic acid. Hydriodic acid is composed of one equivalent of hydrogen and one of iodine; and water, iodine, and hydriodate of manganese result from the mutual reaction of two equivalents of this acid, and one of the peroxyde of manganese. The following diagram gives a more precise view of the nature of the action which takes place in this process.

Before decomposition.		After decomposition.	
125 hydriodic acid	{ hydrogen	1	9 water.
	{ iodine	124	124 iodine.
125 hydriodic acid		125	
44 peroxyde of manganese	{ oxygen	8	
	{ protox. mang.	36	161 hydri. mang.
<hr/> 294 <hr/>		<hr/> 294 <hr/>	<hr/> 294 <hr/>

If, however, an excess of sulphuric acid is present in the liquid, which is generally the case, then only one equivalent of hydriodic acid, being called into action, is decomposed, sulphate of manganese being formed, and the oxygen disengaged from the peroxyde, uniting with the hydrogen of the hydriodic acid, forms water.

When it is difficult to obtain the liquid directed to be employed, a little hydriodic acid may be mixed with the peroxyde of manganese in a small tube or glass retort, to show the action that takes place.

Iodine is a solid substance, of a dark, bluish-black color and metallic lustre; when slowly sublimed, its vapor condenses in rhomboidal plates. It has a pungent odor, an acrid taste, stains the skin of a deep brownish-yellow color, and destroys the vegetable colors, though it acts more feebly upon them than chlorine. With oxygen, hydrogen, nitrogen, sulphur, phosphorus, and the metals, it forms an important class of compounds, similar in their general chemical relations, to those which chlorine forms with the same substances.

Put a few grains of iodine into a large glass flask, and expose it to a gentle heat. The iodine will be sublimed and form a rich, violet-colored vapor, which condenses in crystals, as it cools. Exp.

Boil a little more with five or six ounces of water, in a Florence flask. The vapor of iodine rises along with the vapor of the water, at  $212^{\circ}$ . Exp.

Iodine is called an incombustible body; but it appears to merit the title of a supporter of combustion. If a few grains

be placed in a flask or matrass, and converted into  
 Exp. dense vapors by heat, a small piece of potassium will be inflamed, soon after being thrown in, if a slight degree of heat be applied.

Cut three or four very thin pieces of phosphorus and throw a little iodine upon them after they have been dried, and placed in a tin cup. They combine together, consider-

Exp. able heat is produced, and part of the phosphorus is generally inflamed. If the mixture is made in a small tube, filled with mercury and inverted in a cup of the same liquid, or over the mercurial trough, they immediately combine, but no light is disengaged; it appears only when part of the phosphorus is inflamed by the heat produced, and combines with the oxygen of the air.

Make a strong solution of starch in boiling water, and mix some iodine with it; a compound is formed, immediately, of a very deep blue color. Pour some boiling water upon it, and

Exp. the mixture will become quite colorless; the nature of the changes that take place, on the addition of the hot water, have not been minutely examined. Starch is the most delicate test of iodine which we possess, and it is stated, that, with this substance, one part of iodine may be detected in 450,000 of water; and if we drop a solution of starch into a solution of iodine in water, and observe the deep blue color which is immediately produced, though this liquid can dissolve only one seven-thousandth part of its weight of iodine, we shall not be inclined to doubt the accuracy of the statement.

The best method of detecting minute proportions of iodine, in solution, was pointed out by M. Balard. The liquid suspected to contain it, is mixed with a solution of starch, and sulphuric acid added in excess, after which chlorine water is poured over it; a blue band is perceived where the two liquids

Exp. meet, if any iodine is present, and though it may be very feeble, it is in general distinctly recognised on contrasting it with the liquids above and below. The sulphuric acid separates the hydriodic acid (for this is the form in which iodine usually exists) from any base with which it may be combined, and the chlorine taking the hydrogen, disengages the iodine, which immediately combines with the starch. The blue compound is produced solely by their combination; pure hydriodic acid has no action on starch.

If the proportion of hydriodic acid in solution is not exceedingly small, the sulphuric acid reacts on it and disengages a portion of iodine, which immediately produces the characteristic blue color with the starch, though no chlorine is added.



To see this, dissolve a grain of the hydriodate of potash in a few ounces of water, and pour a solution of starch into it, after adding a drop or two of sulphuric acid. If the solution is strong, iodine vapor is disengaged on adding strong sulphuric acid. Exp.

In all these experiments for detecting iodine, cold water must be employed, as the blue compound of iodine and starch is decomposed by hot water.

Iodine is frequently employed in medicine, both internally and externally; more especially in scrophulous affections. Taken in an overdose, it acts as a virulent poison, even more so than chlorine; and its more common use increases the probability of its fatal results. There is reason to believe, that where it has been given even in small doses, without apparent effect, it has suddenly, as if by accumulation, occasioned very disastrous consequences.\* The only substances which appear to act the part of antidotes, are large quantities of starch and mucilaginous matters of various kinds, dissolved in water.

We shall notice the proper tests of iodine, more fully, under the article hydriodic acid.

### *Of Iodic Acid.*†

Two compounds of iodine and oxygen have been noticed in the books, viz. the *iodous* and *iodic acids*, the first of which appears to be of doubtful existence.

The iodic acid is composed of iodine 124, (one equivalent,) and oxygen 40, (five equivalents,) and is prepared most easily by the process of Sir H. Davy. He directs us to take protoxyde of chlorine (prepared by pouring 400 grains of muriatic acid, specific gravity 1.105, on 100 of the chlorate of potash in a tube retort and applying a gentle heat) and distil it into a small thin glass receiver, containing about 40 grains of iodine, a few pieces of chloride of calcium being put into the neck of the retort. An evolution of heat and light takes place when the protoxyde comes in contact with the iodine, one portion of which, combining with its oxygen, is converted into iodic acid, and the rest unites with the chlorine, forming a compound which is easily separated from the iodic acid by a moderate heat.

It is a white, semi-transparent, solid substance, having a very acid, astringent taste, and is decomposed when heated to 390°. It is soluble in water and deliquescent, reddens and then de-

\* See Journal Complementary, vol. xviii.

† Called by Sir H. Davy, *ox-iodine*.

stroys the vegetable blues. Its salts are termed *iodates*. Many of them may be prepared by the mutual action of their bases on water and iodine, the same reaction taking place that has been already described with respect to chlorine; every six equivalents of iodine decomposing five of water and forming one of iodic and five of hydriodic acid, both of which remain in combination with the oxyde employed. They detonate with inflammable substances, and have the same general properties as the chlorates.

### *Of Hydriodic Acid.*

This, as its name denotes, is a *hydracid*. It consists, by weight, of 124 iodine (one equivalent) and 1 hydrogen (one equivalent); by volume, it is two whole measures. Its specific gravity is 4.337. One hundred cubic inches weigh 132.189 grains. The liquid acid, that is, the gas absorbed by water, has a specific gravity of 1.5.

Hydriodic acid gas is prepared, by mixing one part of phosphorus with ten of iodine moistened with water, and put previously into a very small glass retort or flask, applying a gentle heat with a spirit lamp. In a very short time, a brisk reaction commences, a slight explosion generally taking place within the retort, from the heat produced inflaming a portion of phosphorus, and also from the disengagement of a little phosphuretted hydrogen. Dense vapors are at the same time disengaged, and the hydriodic acid gas may be collected, by displacement, after these have been expelled. Water absorbs it as rapidly as muriatic acid gas, and it cannot be kept long over mercury, as this metal begins to act upon it whenever they come into contact, mercury combining with the iodine, and thus detaching hydrogen gas. A few drops of water should be introduced from time to time by a small dropping tube, as phosphuret of iodine is sublimed into the neck of the retort, when the materials are dry, and no gas is produced. Phosphuretted hydrogen is disengaged in considerable quantity, towards the end of the operation; when it begins to come, it is recognised by the acid gas with which it is mixed, producing a whiter colored vapor than usual, by contact with the air; the process should then be stopped, to prevent it from accumulating. Constant attention must be paid to this operation while it is going on.

A number of complicated changes takes place during the preparation of this gas, from the reaction of the different substances mixed together and part of the newly formed products. The hydriodic acid gas is produced, by the iodine combining with the hydrogen of a portion of water which is decomposed,

the oxygen going to the phosphorus and forming phosphoric acid, which remains in the retort and acts upon the glass, if the heat is urged after it has been formed. Supposing the reaction to take place between two equivalents of iodine, two of water, and one of phosphorus, and no other compounds to be produced, the following diagram will convey a clearer view of the manner in which the different substances arrange themselves.

Iodine	124	125 hydriodic acid.
Iodine	124	125 hydriodic acid.
Water {hydrogen	1	
{oxygen	8	
Water {hydrogen	1	
{oxygen	8	
Phosphorus	12	28 phosphoric acid.
	<hr/> 278 <hr/>	<hr/> 278 <hr/>

Fifty or one hundred grains of iodine, with the proper quantity of phosphorus, will be found quite sufficient, using a retort capable of containing about five or six ounces of water.

When a solution of hydriodic acid in water is required, in which state it is usually kept as a test, the best process is to decompose the iodide of starch, suspended in water, by a stream of sulphuretted hydrogen. (Journal of Science, New Series, No. viii.) Sixty grains of iodine are dissolved in three ounces of alcohol, (in the cold,) and an ounce of starch reduced to a very fine powder, diffused in four ounces of water, adding this, drop by drop, to the first solution, and stirring it constantly at the same time; iodide of starch is formed, and the clear liquid is decanted after it has subsided. A little water is then poured on it to remove any alcohol that may still be mixed with it, and after it has been removed, the iodide is diffused through an ounce of water, and a stream of sulphuretted hydrogen (from 400 or 500 grains of the sulphuret of iron) passed through it till it becomes white. The liquid is then filtered to remove the starch and sulphur that are disengaged, and boiled for a short time to expel any excess of sulphuretted hydrogen. The iodide of starch may be put into a common precipitate glass or jar, when it is diffused through water, and the sulphuretted hydrogen prepared in a bottle with a bent tube fitted to it, or any of the other forms of apparatus that may be preferred.

In this process, the hydrogen, of the sulphuretted hydrogen, combines with the iodine of the iodide of starch and forms the hydriodic acid, which remains in solution, the sulphur and

starch being separated; none of the starch is taken up by the solution, as it is insoluble in cold water. The mixture passes through a variety of shades of color, from the deep blue of the iodide, to a rich brownish-red, orange, and yellow color, before it becomes ultimately white; these changes succeed each other rapidly, and present a very beautiful appearance when the sulphuretted hydrogen is quickly evolved.

A solution of hydriodic acid in water may be obtained, also, by transmitting sulphuretted hydrogen through iodine reduced to a fine powder and suspended in water, the hydrogen combining with the iodine, and the sulphur being deposited. I prefer the process already described, as the iodine is obtained in a much more minute state of division than it can be procured by trituration, and the sulphuretted hydrogen acts more readily upon it in this state. In both cases, the liquid acid may be concentrated, by evaporation, till it is obtained of the specific gravity of 1.5, continuing the application of the heat for this purpose, till its boiling point rises to  $260^{\circ}$  or  $262^{\circ}$ .

In this state, it is a transparent and colorless liquid, having very strong acid powers and effervescing with carbonates. It acquires a deep color on exposure to air and light, from the decomposition of a minute portion of acid and the separation of iodine.

Exp. Pour a small quantity of the acid into a solution of litmus; the color immediately changes to a red.

Pour a few drops into seven glasses, each containing an ounce or two of water. To the first, add a single drop of a solution of the chloride of platina; the whole liquid will

Exp. immediately become of as deep a reddish-brown color as the strong solution of the chloride employed, and after some time a dark, brownish-black precipitate is formed. The chloride of platina is the most delicate test of hydriodic acid which we possess.

Into the other glasses pour a few drops of the following liquids.

1. A solution of the nitrate of silver; a yellow precipitate immediately appears.

2. A solution of the bi-chloride of mercury gives a precipitate which appears yellow at first, but soon becomes of a brick-red color.

3. A solution of perntrate of mercury gives a similar precipitate.

4. A solution of the acetate of lead gives a yellow precipitate. All these precipitates are compounds of the iodine of the acid with the metal of the solution employed.

5. Strong nitric or sulphuric acid decomposes it.

6. Chlorine water produces the same effect, the hydrogen being withdrawn and the iodine set at liberty; a solution of starch may be added afterwards, to produce the characteristic blue precipitate.

With the hydriodic acid gas, collected by displacement, several experiments may be performed. If any has been collected over mercury, it should be used as soon as it is prepared, leaving one jar to show its complete decomposition by this metal, and, also, that half its bulk of hydrogen remains.

Remove a tube filled with the gas, and take the finger off the mouth under water; the gas will be absorbed as rapidly as muriatic acid gas, if it is perfectly pure. Exp.

Introduce a small quantity of water into a jar full of the gas, over mercury, to show the large quantity which it can absorb. Exp.

Fill a small jar half full of chlorine over the mercurial trough, and transfer it immediately, under the mercury, to another jar half full of hydriodic acid gas; the chlorine combines immediately with the hydrogen of the hydriodic acid, forming muriatic acid gas, and the iodine is deposited. Exp.

Invert a jar full of hydriodic acid gas with a small glass plate, and pour into it a little of the strong fuming acid, composed of nitric and nitrous acids. The hydrogen of the hydriodic acid immediately combines with the oxygen of the fuming acid, and iodine is set at liberty; the mixture often inflames, even when the experiment has been made with no more than two or three cubic inches of the gas.\* Exp.

### *Of Iodide of Nitrogen, Chloriodic Acid, Cyanide of Iodine, &c.*

*Iodide of nitrogen* is prepared by triturating iodine with liquid ammonia, allowing the mixture to remain for two or three hours. Ammonia is composed of hydrogen and nitrogen; part of the iodine combines with the nitrogen, forming the iodide, which assumes the form of a black powder, and another portion, uniting with the hydrogen, is converted into hydriodic acid, which remains in combination with the water of the ammonia. It detonates by a slight pressure, or on exposing it to a moderate heat over a chauffer, its elements being

\* Hydriodic acid, in combination with potash or soda, is found in several mineral waters in Europe and in this country. Possibly, the utility of these waters, in scrophulous diseases, may be owing to the presence of iodine, in this form.



separated from one another. It may be prepared in a small evaporating dish, and detonated in the same vessel, after pouring off the solution of hydriodic acid. A few grains of iodine and a drachm or two of ammonia will afford a sufficient quantity to show its detonating properties.

The action that takes place between iodine and phosphorus has been already described. A small quantity of the iodide of phosphorus is generally sublimed during the preparation of hydriodic acid gas, condensing in the neck of the retort; if a little water is poured upon it, an immediate effervescence takes place, and a large quantity of hydriodic acid gas is produced, the iodide reacting on the water, in the same manner as the mixture of iodine and phosphorus.

*Iodide of sulphur* is prepared by mixing iodine and sulphur together and exposing them to a gentle heat.

*Chloriodic acid* or *chloride of iodine* is the name given to a compound of chlorine and iodine discovered by Sir H. Davy. It is prepared by admitting chlorine into an exhausted flask, containing a fixed quantity of iodine. It has not been applied to any use.

*Cyanide of iodine* is a compound of cyanogen and iodine, made by mixing two parts of bi-cyanide of mercury with one of iodine, intimately and quickly in a glass mortar, and putting the mixture into a wide-mouthed vial. On applying heat, vapors of iodine appear, but when the bi-cyanide is decomposed, its cyanogen joins the iodine, and the violet vapors are succeeded by white fumes, which soon condense into flocks, like fine cotton. It is an unimportant article.

---

## CHAPTER XI.

### Of Bromine.

THE French usually call this substance *brome*, and this name is sufficiently correct; it is derived from a Greek word, signifying a fetid smell. Its equivalent is 75.

*Bromine* was discovered only a few years ago, by M. Balard; it is obtained by passing a stream of chlorine through bittern, (the liquid that remains after boiling down sea-water to prepare common salt,) and exposing it afterwards to heat, when bromine is distilled, and may be collected in a receiver.

Another method is to shake sulphuric ether with the bittern,

after the chlorine has been passed through it; the ether dissolves bromine, and when it is left at rest, it collects at the top, having a very rich hyacinthine color. On digesting it with a strong solution of potash, two new salts are formed, the bromate and hydrobromate of potash; the latter is obtained in cubical crystals by evaporation, and on mixing it with peroxyde of manganese and sulphuric acid in a glass retort, the bromine is disengaged by the application of heat.

Bittern consists principally of sulphates and muriates of soda and magnesia, with a small quantity of the hydrobromate of magnesia, a compound of hydrobromic acid and magnesia. The hydrobromic acid is composed of hydrogen and bromine, and when the chlorine is transmitted through the bittern, it combines with the hydrogen and disengages the bromine, which imparts a yellow color to the liquid. The vapor of the bromine has a deep, reddish-brown color, bearing a great resemblance to nitrous acid vapor, and condenses into a very dark colored liquid.

In the second process, the bromate and hydrobromate of potash are formed, by one portion of the bromine combining with the oxygen, and another with the hydrogen of a portion of water which is decomposed. When the hydrobromate is afterwards mixed with sulphuric acid and peroxyde of manganese, sulphates of potash and manganese are formed, oxygen gas being disengaged and taking hydrogen from the hydrobromic acid, which is eliminated at the same time, and the bromine, which is set at liberty, is collected in the receiver.

A few ounces of bittern are sufficient to show the first process for the preparation of bromine, though little more than a grain or two is obtained; but for the second, a larger portion must be employed. In preparing the liquid, the chlorine must be transmitted through the bittern till the orange-yellow color which it acquires, ceases to become deeper.

Bromine is the only elementary substance that is liquid at natural temperatures, with the exception of mercury. Its present appellation is derived from its disagreeable smell, that being highly characteristic. In all its chemical relations, it bears a great resemblance to chlorine and iodine, and its compounds are analogous to those which they form with the same bases; they may be prepared in the same way. For a more particular account of its properties, I must refer to the different scientific journals.

Bromine combines with the metals, forming *bromides*. It unites, also, with the simple, non-metallic bodies, giving rise to compounds, that have not any practical importance.

*Bromide of chlorine*, or *chloride of bromine*, is made by passing chlorine through bromine, and condensing the volatile vapor, when cold. It is a reddish-yellow liquid, very fluid and volatile, soluble in water, and possessing bleaching powers. Metals burn in its vapor.

*Bromide of iodine* is formed by the direct union of the respective vapors of the two components. It is a solid compound, which yields, by heat, brownish-red vapors.

*Bromide of phosphorus*. Bromine and phosphorus readily unite, even in a flask containing carbonic acid gas. Light and heat are emitted, a crystalline solid being sublimed, while a liquid subsides. The latter is a protobromide, the former a deutobromide.

*Bromide of sulphur*. This may be formed by adding bromine to flowers of sulphur. It is a dense liquid, of a deeper color than the chloride of sulphur, and like it, diffuses white vapors.

*Bromide of cyanogen*, or *cyanide of bromine*, is prepared by a process similar to that for making cyanide of iodine. The two compounds bear a close resemblance to each other, but the former is more volatile than the latter. It is highly deleterious to animal life, as shown by experiments made on rabbits.

*Bromic ether* is formed by letting fall a drop or more of bromine into a flask of olefiant gas. An oily substance, heavier than water, is produced, destitute of the odor of bromine, and having a more agreeable smell than chloric ether.

Bromine exists in many mineral waters, but always in the form of a hydrobromate. It has been detected in several of the springs in this country.

---

## CHAPTER XII.

### Of Fluorine.

THE existence of this simple, non-metallic substance is a matter of inference, rather than demonstration. It is now admitted, however, by all chemists, and is the elementary basis of fluoric acid and the fluuate of lime, known as *Derbyshire* or *fluor spar*.

*Fluoric acid* is a colorless, transparent fluid, emitting copious, dense, white fumes, on exposure to the air. Its specific gravity is 1.06. It is generally regarded as a compound of fluorine 16, (one equivalent,) and hydrogen 1, making its equivalent 17.

To prepare this acid, we decompose the fluor or Derbyshire spar, (a compound of fluoric acid and lime,) by means of sulphuric acid. The spar must be reduced to a very fine powder, and be mixed with its own weight of sulphuric acid, by means of an iron rod. As this acid acts powerfully on glass, and is thus changed into another substance, it is necessary to employ metallic vessels for its preparation. Retorts and receivers of lead are generally used, as the fluoric acid has no action on that metal; but any leaden vessel, of a hollow structure, will answer sufficiently well. On applying a moderate lamp heat, decomposition is fully effected, and the fluoric acid is disengaged, while sulphate of lime remains in the vessel. If the heat be too great, the lead will be melted. If a receiver be used, in connexion with the retort, it should be kept cold by means of ice or cold water.

In forming the acid, to show its etching properties, I have employed a leaden box of six inches square, and two inches deep. In this, about two ounces of the powdered spar may be placed, and the same quantity of acid. If a sand bath be at hand, duly heated, it will answer better than a lamp or chauffer, and does not endanger the lead in the least. The glass to be etched, should be a little larger than the leaden box; and having been coated with wax, and the figure drawn, it should be laid closely on the box, and a weight applied to keep it so. In a few hours, the acid will have acted on the glass, so as to make an impression of the entire figure, drawn on the plate, and on removing the wax, and washing with a little spirits of turpentine, the etching will be apparent.

Fluoric acid has all the properties of a very powerful acid, and is particularly distinguished by its action, as above stated. It acts on the silica of the glass, and forms a gaseous compound, to be noticed presently, called fluosilicic acid.

The greatest caution is necessary in experimenting with this acid, as its vapors are not only highly irritating, but it acts, also, more violently on animal substances, than any other acid. If a single drop fall on the skin, it is instantly destroyed, and a deep ulceration follows, that is not easily cured.

When the diluted acid is used for etching, instead of the acid vapor, it may be poured on the coated glass, after the figures have been traced on it, taking care to have a rim of wax on the border of the plate, sufficiently high to prevent the acid from running off. The liquid acid is preferable to the vapor, as it does not impair the transparency of the glass.

When the strong acid is dropped into water, a hissing noise follows, attended with much heat, and the diluted fluid is capable of acting on glass.

Until the real nature of fluoric acid be well ascertained, we cannot give a satisfactory theory of the reaction that ensues, in the process for forming it. Many persons regard the fluor spar as a compound of dry fluoric acid and lime, and according to this view, the dry sulphuric acid (in the liquid acid employed) combines with the lime, and the fluoric acid uniting with the water previously in combination with the sulphuric acid, forms the liquid that is condensed in the receiver. The more common opinion, however, is, that fluor spar is composed of fluorine and calcium, (the metallic base of lime,) and that the water of the sulphuric acid is decomposed, the oxygen combining with calcium and forming lime, which remains in combination with the sulphuric acid in the leaden vessel, while the fluoric acid is formed by the fluorine combining with the hydrogen.

Gay Lussac and Thenard, who first obtained this acid in a pure state, (in 1810) regarded it as a compound of a certain inflammable principle and oxygen, and they called it *fluoric acid*.\* Those who consider it as a *hydracid*, or an acid the base of which is acidified by hydrogen, call it *hydrofluoric acid*. The salts, agreeably to the views of some, are *fluates*; according to others, *hydrofluates*. If fluor spar be a compound of fluoric acid and lime, then the views of Gay Lussac are correct; but if it be a fluoride of calcium, (a compound of fluorine and the metallic base of lime,) then it follows that the more general opinion is the one that is entitled to credit. The experiments of Mr Kuhlman seem altogether conclusive. He found that fluor spar could not be decomposed, in the slightest degree, by means of anhydrous sulphuric acid, either at common temperatures or at a red heat. If the production of fluoric acid depended on the simple decomposition of the fluor spar, this should be effected by the strongest sulphuric acid. But as no decomposition occurred in the experiments of Kuhlman,† the inference is favorable to the idea, that water must be present, in order to furnish hydrogen by its decomposition, which, uniting with the fluorine of the spar, gives rise to the hydrofluoric acid.

*Fluoboric acid*. This is a compound of fluoric and boracic acids, and is made by exposing boracic acid and fluor spar, in a tube, to a red heat. It has been called, also, *borofluoric* and *fluoboracic acid*; either name will answer our purpose. The easiest method for making this acid, is to mix, intimately, fused boracic acid with twice its weight of fluor spar, (both having been first pul-

\* The name given by its discoverer, Scheele.

† Quarterly Journal of Science, July, 1827.



verised,) and twelve parts of sulphuric acid. The mixture should be made in a glass retort, and heated by a chauffer or spirit lamp. An ounce of the fused boracic acid, with corresponding portions of fluor spar and sulphuric acid, afford a considerable quantity of the compound.

Fluoboracic acid exists in the gaseous state, and is largely absorbed by water, which condenses 700 times its volume, and acquires a specific gravity of 1.77. So great is its affinity for water, that it abstracts it from all gases containing it, forming very dense, white fumes when allowed to mix with atmospheric air.

*Fluosilicic acid gas.* This compound has been already adverted to, as resulting from the action of fluoric acid on glass. The silica of the glass combines with the acid of the fluor spar, and gives rise to the compound now under examination. It is a permanent gas at the ordinary temperature and pressure.

To make it largely, take one part of pounded glass, an equal quantity of fine dust of fluor spar, and two parts of sulphuric acid, and mix them well together. A slight degree of heat accelerates the process, which should be conducted near the mercurial trough, so that the acid gas may be collected over mercury. It cannot be obtained over water, because that fluid decomposes it, occasioning the precipitation of one third of the silica, while the whole of the acid gas and the residue of the silica, are dissolved by the water, which takes up 365 times its bulk of the acid gas.

Fluosilicic acid gas is exceedingly pungent and irritating; it does not corrode dry glass vessels, is not a supporter of combustion, and forms white fumes with air that contains watery vapor. Its composition is necessarily involved in some degree of obscurity; being regarded as a compound of fluorine and silicum, (the metallic base of silica,) on the one hand, while, on the other, it is supposed to be a compound of fluoric acid and silica.

---

## CHAPTER XIII.

### Of Combustion.

THIS term, in its common acceptation, means an actual burning, in which heat and light are evolved, and the bodies burned are actually changed. It differs from simple ignition, inasmuch as the latter merely indicates the glowing state of a body, said

to be *red hot*, independently of any real chemical change. *Ignition* and *incandescence* are synonymous terms.

According to Sir H. Davy, combustion may be an accompaniment of intense chemical action, no matter what be the nature of the substances present, whether combustible or incombustible, supporters or non-supporters, according to the usual acceptation of those terms.

It is not my intention to give an elaborate view of this very interesting topic, but briefly to notice three important theories or views, which have prevailed at different periods. These are the *Stahlian* or *phlogistic*, the *Lavoisierian* or *anti-phlogistic*, and the *chloridic* theories.

For many years before the discovery of oxygen gas was announced, all the phenomena of combustion were explained on the *Stahlian* or *phlogistic* hypothesis. It was supposed by Stahl, that all bodies, susceptible of being burned, necessarily contained within their texture, a certain, peculiar principle, to which he gave the name of *phlogiston*, a word whose Greek derivative means, to burn. This *phlogiston* was the inflammable principle, without which, combustion was impossible. When a body was burning, Stahl supposed that its *phlogiston* was escaping, and that so soon as the whole of this inflammable principle had passed out, the body was *dephlogisticated* or *incombustible*. The heat and light of combustion, were ascribed to the rapidity or slowness, with which this principle was evolved. All the bodies that are now called simple, inflammable substances, were regarded as compounds, of which *phlogiston* made an essential part, and the products resulting from their combustion were called simple, because, in burning, they lost the *phlogiston*.

An important error in this system, was the notion that bodies became lighter, in the act of burning; and this error resulted from false theory, apart from experiment or correct observation. Affirming that *phlogiston* was a substantial reality, and that it escaped during the combustion of bodies, there was no alternative but to conclude, that bodies were made lighter by the loss of this imaginary principle.

It is easy to perceive, that such a system could not long maintain its ground, when attacked by the powers of demonstration. The discovery of oxygen and the announcement of its properties, prepared the way for the overthrow of the *Stahlian* doctrine. It was reserved for the inimitable Lavoisier to expose the fallacy of this system, and to substitute in its room, the *anti-phlogistic* theory. Agreeably to the accurate experiments of this chemist, combustion and oxydation consisted in the combination of the combustible body with oxygen, and this, as a general rule, he

succeeded in establishing, beyond a doubt. I care not for the national jealousy and envy which leads men of the high standing of Dr Ure, to attempt to depreciate the value of the anti-phlogistic system, and to tear from the head of Lavoisier, his well-earned laurels. The time will yet come, when chemists will be willing to do full justice to the labors of that indefatigable but unfortunate man, and when the scientific fruits of his short life will be deemed vastly more important, than the entire mass of mystical philosophy of a score of his opponents. I maintain now, as I have done elsewhere, that the spirit of the anti-phlogistic system is the true spirit of chemical science, and will continue to be so, as long as the inductive philosophy shall have place in the earth. There is not a feature in all the modern discoveries, which cannot be associated and incorporated with that spirit, without risk of deformity or injury.

But we have said, that Lavoisier overturned the Stahlia doctrine. He did so, by ascertaining that combustibles, in the act of burning, did not become lighter, but heavier; that instead of parting with something, they actually combined with oxygen, and that the increase in weight of the body burnt, exactly corresponded with the quantity of oxygen that was consumed. For instance, he burned iron wire in a vessel full of oxygen gas, and he found that the oxyde, thus formed, weighed precisely as much as the original wire and oxygen that were present prior to the combustion. That the oxygen was really present in the oxydised body, he proved, by exposing mercury to heat in the open air, until it absorbed as much oxygen as changed it into an oxyde; after which he placed the mercurial oxyde, thus formed, in a retort heated to redness, and recovered the mercury and oxygen in their original forms and quantities.

Lavoisier probably erred, in accounting for the heat and light evolved in combustion, by reason of the limited view he took of the subject. It would require an infinite number of experiments, to determine with precision the source of light and heat in all cases of combustion, and the precise quantity of these evolved, in each instance, from each body submitted to the process. We say, by way of general remark, that light and heat are properties of all sorts of matter, and that when these are evolved during combustion, it does not arise from any new creation, but from their simple disengagement from the burning bodies, in which they exist by nature, as inherent attributes.

If discovery, in regard to supporters of combustion, had ceased with the production of oxygen, the Lavoisierian system had retained its popularity to the present hour. But the inquiring curiosity of the chemist was not satisfied with present attain-

ments, and his diligent researches soon brought to light, chlorine, iodine, and bromine. In the honor of these discoveries Sir Humphrey Davy had the largest share, and the world is deeply indebted to him for the fruits of his untiring zeal.

The announcement of chlorine and its striking peculiarities, constituted the reasons for what has been designated, *the chloridic theory*. As *oxymuriatic acid*, this substance had been familiar, as a supporter of combustion; but its powers in that character were ascribed to the presence of oxygen, its compound nature being then universally acknowledged. And such was the tenacity with which long-cherished opinions were held, that the compound quality of oxymuriatic acid was contended for, long after the publication of experiments proving the elementary nature of that body. The most able opponent of the simple character of oxymuriatic acid, was Murray, and for a while he enlisted many on his side of the question. My own inaugural dissertation\* was, in part, an attempt to disprove the simple character of chlorine, but I am happy to acknowledge, that so far as experiments appear to go, we are bound to regard chlorine as an element.

In this state of things, it ceased to be true, that oxygen was the only simple substance possessing the attribute of supporting combustion. But chlorine did not long stand alone, but was followed by iodine, bromine, sulphuretted hydrogen, and others, which, under given circumstances, evidently maintain the character of supporters of combustion.

These apparent exceptions to the spirit of the Lavoisierian system, seem rather to accord with it, or to be extensions of it. All of them are at war with the notion of the phlogistic system, that combustibles, in burning, become lighter. And even in their bearings on the doctrine of acidification, they present analogies, rather than contrarieties, to the anti-phlogistic system.

In conclusion, we regard combustion as a result of the action of material particles upon each other, as influenced by strong chemical attractions, or opposite electrical conditions, and that it is a necessary consequence of an intense and violent motion between the molecules of matter, whatever be the exciting cause. As light and heat are inherent properties of all matter, we hold, that the evolution of these is not a new product, growing out of the act of combustion, but a mere disengagement or disturbance of these essential constituents of material particles.

\* Dated, March, 1812.

The reader is now prepared to understand the reason for deferring the consideration of combustion to this part of the work. In many treatises, it is introduced as an appendage to the remarks on oxygen or hydrogen; and, in some, it is postponed to the closing pages. As several supporters of combustion were to be examined, each of which was entitled to that appellation, as truly as oxygen, the interest of the student, as well as the consistency of science, seemed to require the adoption of the course here pursued.

This is not the place for details of facts or arguments on these controverted points, which cover a large portion of the more extensive works on chemistry. To these, the leisure of the student may be profitably directed, even though the investigation should fail to rescue the subject from doubt and obscurity.



## CLASS SECOND.

OF METALS AND THEIR COMBINATIONS WITH NON-METALLIC  
SUBSTANCES, AND WITH EACH OTHER.

---

### GENERAL REMARKS.

HAVING examined those bodies which are usually denominated *simple non-metallic*, we are naturally led to investigate the metals, or the *simple metallic bodies*. But before we enter on the history of each particular metal, we think it indispensable to say something of the general properties of those substances, and which characterise and distinguish them from all other bodies. These things having been set forth in this part of the work, we shall not be under the necessity of repeating them, when considering each metal by itself.

We shall, in the first place, mention in what state metals are found in nature; we shall point out the preliminary preparations necessary to be attended to, in order to separate them from those substances with which they are mechanically mixed, and after having presented some thoughts on the manner in which they are brought to a state of purity, we shall examine the action exerted upon them by air, oxygen, water, &c., and we shall show what agency all these considerations have had, in reducing the metals to a methodical classification.

Metals are found in the bowels of the earth, as well as at its surface; they are found, more especially, in what are called *primitive earths*, and in earths of transition, under the form of masses or veins. The first of these arrangements is the most frequent; the metal seems to have been thrown in a heap, in one spot, and as if by chance; the second, viz. the vein formation, consists of metallic layers of different thicknesses, (sometimes a few inches, and now and then of as many feet,) sometimes placed horizontally, more often obliquely, and but seldom in a vertical position. These masses or veins, as they are distributed over an extent of earth more or less considerable,

constitute a *mine*. Mines are more commonly found in mountains, and especially in those of a mean altitude, in the formation of a ridge.

When a mine has been discovered, the first effort is, to extract the metal. Metallic substances are almost always combined with foreign matters, which affect their purity; they are placed between two stony layers or strata, of which the superior one is called the *roof*, because it affords a covering, the inferior one being called the *bed*, because the metal rests upon it. These two strata constitute what is called the *gangue*; they vary in solidity, as well as in thickness, and are often composed of quartz, barytes, fluor spar, &c.

It forms no part of the present design, to describe the mechanical processes usually employed for the extraction of the mineral; these details belong to the art of *mining*. It will suffice to say, that in order to reach the mine, it is necessary to make what are called wells, and this is done by boring; these are of various depths, according to the course of the metallic strata. They are made to communicate with each other in the oblique direction of the veins. This disposition of the wells is not only favorable to the exploration of the mine, but is highly advantageous, by establishing currents of air which are well calculated to drive out the exhalations from these subterranean excavations, and which very often prove hurtful to the miners.

Previous to working the mine, we should ascertain whether it will be profitable, or merely meet the enormous expense of the undertaking; and to this end, an assay of the mineral should be first made. This is a business which requires much care and exactness; for it will not suffice to extract the mineral from foreign matters; it is indispensable to know, with something like certainty, what proportion the metal bears to the entire mass. In order to arrive at a correct result, it is obvious that specimens should be selected of a mean quality, taking neither the best nor the worst. Much embarrassment has often arisen from a wrong selection; the examination of rich specimens has raised false and delusive expectations, and ultimately ruined the speculator. When the processes of the assayer have determined the proportion of metal contained in the mineral, a judgment may be formed with propriety, on the question of working the mine. The mineral is first to be separated from its *gangue*, by mechanical means. The first operation is called distribution. This consists in separating the pieces which appear to contain a considerable quantity of metal, from those which would not be likely to yield a sufficient product. This being accomplished, the mineral must be pulverised in a wooden

mortar and with a wooden pestle, whose inferior extremity is armed with iron. Next, the earthy parts are to be separated by washings. This process is conducted on tables, or in large troughs, by means of streams of water that extract the lighter particles, which are brought to the surface by frequent agitation of the mixture. To this operation follows the *roasting*, which consists in exposing the metal to a very strong heat, in order to disengage the volatile substances which are combined with it. These are, usually, sulphur and arsenic, which the ancient chemists styled the *grand mineralisers*, because very few metals are found free from this admixture. The *roasting* accomplishes another end, viz. the separation of the water, (which, in a measure, binds the substances together,) to destroy the force of cohesion in the mineral, and to make the subsequent separation of the metal, by means of solvents, more easy.

We come, now, to notice the manner in which metals exist in nature. They are presented to our view, under four principal aspects.

1. The *native* or *virgin* state, by which we do not mean a condition of perfect purity, for they are almost always more or less affected by foreign substances, but we consider them *native*, when their exterior characters are so obvious, that an experienced mineralogist can easily recognise them. The metals which are the least altered in appearance by mixture, as might easily be supposed, are those which are found in the native state, as gold, platina, silver, &c.

2. The metals which combine readily with oxygen, or are easily oxydated, as copper and iron, are rarely found in the native state, but usually in the state of oxydes. There are some, as manganese, whose affinity for oxygen is so powerful, that we scarcely ever find them in any other state.

3. A third condition is that of the natural combination of metals with non-metallic substances. It is thus that sulphur is combined with iron, copper, antimony, silver, mercury, zinc, and forms with them native sulphurets, in great abundance. Chlorine is found united with the metals, especially to mercury, silver, and, occasionally, to lead. Carbon is found combined with iron, in the form of natural steel, in the lead mines of England, as a carburet of iron, (graphite and plumbago). M. Vauquelin, a few years since, discovered the combination of iodine with silver, in a Peruvian mine; and selenium has also been detected, in combination with silver, copper, and iron, in the form of seleniurets.

4. The fourth condition in which we find metals, is the *saline state*. We have said, that many of the metallic substances are combined with oxygen; so that it will not be difficult to imagine,

that an oxydated metal, coming in contact with acids, may combine with them, and give rise to salts. Sulphuric, carbonic, phosphoric, and fluoric acids, are those which are most frequently combined with the metals. There are some metals, as arsenic, which have the property of combining with metallic oxydes; hence we have arseniates of copper, lead, iron, &c. Molybdena, tungsten, chrome, and columbium have the same property, and form, also, in combining with oxydes, salts which are known under the names of molybdate of lead, tungstates of lime and iron, columbate of iron, chromate of lead, &c.

Such are the four principal conditions, in which we find metals in nature. It is the province of metallurgy, or the art of managing metals in the large way, to separate the foreign matters that are mixed or combined with them, and to bring them to that state of purity, which their use in the arts demands. The metallurgic processes are very numerous, and may be varied according to their application to this or that metal. They are divided into mechanical or preparatory processes, such as *distribution, pulverising, washing, roasting*, as we have already noticed; and into chemical processes, strictly so called; of these latter we shall speak, whenever a metal of much importance is under consideration. We shall content ourselves, at present, to select from the more simple operations, one or two examples, which may serve to give an idea of the method by which metals are purified.

When we wish to obtain pure antimony, we pulverise the sulphuret of that metal and then roast it. A large quantity of sulphur is disengaged during the operation; but however long we may continue the roasting, or however great our precautions, in other respects, may be, more or less sulphur will remain. Further; in proportion as the sulphur is detached from the antimony in the state of sulphurous acid, the metal undergoes a change, and is oxydated, so that after the roasting, we have an oxyde of antimony mixed with sulphur; that is to say, that the antimony is, really, less pure than it was before; we have then to separate it, both from the sulphur which is hidden, and from the oxygen which has been absorbed. To this end, we must add to it a mixture of carbonate of potash and charcoal, known long ago under the name of black flux, and which possesses all the properties necessary for bringing the oxyde of antimony mixed with sulphur, to the state of purity. In the arts, in place of the black flux, coarse tartar is used, because more economical. We take two or three parts of either the black flux or the tartar, for one part of antimony; we make an exact mixture of these two substances, and place them in a crucible; we cause them to melt by means

of a proper heat, and at the end of a half hour, the operation is at an end. We have already noticed the affinity of charcoal for oxygen, and that of sulphur for potash is not less; hence we can imagine how the black flux purifies antimony. If we employ the tartar, there is a disengagement of carbonic acid or the gaseous oxyde of carbon; water and sulphuret of potash being also formed. The antimony, by reason of its weight, remains at the bottom of the crucible; if we cover it with very fine sulphur, so as to exclude the external air, it will be preserved, free from oxydation.

Another method may be employed for purifying the sulphurets, and, especially, the sulphuret of lead; that is, to add a metal which has a stronger affinity for sulphur, than the metal we wish to obtain; for example, if we melt the sulphuret of lead, (galena,) after having mixed it with the third of its weight of iron filings, the pure lead, being more weighty, is deposited; the sulphuret of iron, being lighter, remains at the surface of the mixture. We draw the crucible from the fire, when the matter is in the state of fusion, and suffer it to cool; then, with a sudden blow of a hammer on the mass, we separate the sulphuret of iron from the dross of the lead. This process is successfully employed in the assay of galenas, and for the determination of the quantity of lead, and at the same time of silver, which is afterwards separated, by cupellation.

When the metals are in a state of purity, it is easy to study their physical or exterior characters; some of these belong to almost all the metals; such are the characters by which they are classed together; others appertain only to a certain number of the metals, and serve to combine them in suitable groups to form the divisions of the classes; other characters are purely individual or specific, and fitted to distinguish them from each other.

The characters which may be regarded as appertaining to all the metals, are brilliancy or lustre, opacity, and specific gravity or density. Their lustre has relation, or is in proportion to their faculty of reflecting light; we perceive this in the particles of metals that have been reduced to powder; they differ, in that respect, from certain minerals which, by their shining aspect, might at first rival the metals, but which, in the state of powder, instead of being brilliant, are earthy and dull, in appearance and in reality.

Mineralogists say, that a body is opaque, which, in the thinnest state to which it can be reduced, will not suffer the light to pass through it. The metals possess in an eminent, we might say, in an exclusive degree, this property; indeed, it has been



remarked, that a leaf of gold, made so thin as to be blown away by the breath, if placed between the eye and the light, causes a greenish color to be perceived; but it seems to us, even admitting that in this state of extreme tenuity the leaf of gold does not experience a solution of continuity, we could not thence decide against opacity, as characteristic of metals.

As to specific weight or density, we are aware, that, of all bodies, metals possess it in the highest degree; the least dense weigh six times more than water, whilst earthy or saline substances of the greatest density, those for instance whose bases are metallic oxydes, as the native sulphates and carbonates of barytes and strontian, weigh four times as much as that liquid. To this characteristic of the metals, there are, nevertheless, some remarkable exceptions. The density of sodium and potassium, far from being similar to that of other metals, is inferior to that of water. Should we infer from these exceptions, that density is not a characteristic of the metals? Surely not, since it belongs obviously to almost the whole of these bodies. Where is the class, what is the order or the genus, all the individuals of which exhibit the principal characteristic, without exception.

Independently of these three marks which we have considered as characteristic, because possessed by all the metals, there are other properties which belong only to a certain number, and of which the others are altogether deprived; such, for example, is ductility, or the faculty which belongs to certain metals, of being greatly extended when forcibly compressed. Many of the metals are of this kind, as the greater part of those that were known to the ancients. There is a still larger number, however, which, instead of being much extended, are broken and reduced to powder by percussion, as, for example, bismuth, antimony, and arsenic. These latter are named *brittle*, the others *ductile* or *malleable*. These expressions are not synonymous, although some metals evince both ductility and malleability. It does not necessarily follow, that a metal must be ductile, because it is malleable; *malleability* is the faculty of being extended by the stroke of a hammer, as when a leaden bullet is pounded on an anvil, until it is reduced to the thickness of paper, in which case it occupies a much larger surface than formerly; *ductility* is the faculty of being drawn out into a thin wire, by means of the instrument called the *wire drawer*. This difference in metals is attributed, in general, to the structure of their tissue, whether that be lamellated or fibrous. Lead is very easily reduced to sheets, but is ductile to a very small extent; iron, on the other hand, is very ductile, and is reduced, without great difficulty, to the state of sheet iron. Some metals,

as zinc and nickel, deprived almost entirely of malleability, in the ordinary way, and especially of ductility, are very easily brought to the state of thin plates, by the pressure of the *flat-tener*, as it is termed.

We should observe, that metals do not fully evince their ductile and malleable properties, unless heated previously to subjecting them to the action of the hammer or the wire drawer. Caloric diminishes their force of cohesion, separates their particles, and allows them the more easily to slip or pass by each other.

The malleable metals and, more particularly, the ductile metals, possess a property to which the name of *tenacity* has been given; it is estimated by the weight which a metallic wire, of a given length and thickness, can support without being broken; and it is dependent, most likely, on the adhesion, or the force of cohesion subsisting between the molecules of the metal; but it is not requisite that this property should be alike in all the metals. In some, it is very feeble, as in lead, while others evince it in an eminent degree; thus an iron wire, whose diameter is only one-eighth of a line, (little more than the 100th part of an inch,) will support a weight equal to 540 pounds. Next, in this respect, to iron, is copper; then platina, silver, and gold.

The metals vary considerably in point of hardness. Some are very hard, as tungsten, palladium, iron, copper, &c. Silver and gold are not so hard; tin and lead are still softer, so that we may mark them with the finger nail; sodium and potassium are quite soft, having the consistency of wax; and mercury is liquid. The hardness as well as the elasticity of iron may be augmented, by combining it with a small quantity of charcoal, which converts it into steel. We increase the hardness of copper, and at the same time, its sonorousness, by alloying it with tin; this is what constitutes bell-metal. Copper, alloyed with gold and silver, communicates to these metals a solidity, without which, the manufacturers of fine jewelry could not work them.

A few of the metals are odorous and sapid. Iron, copper, lead, and tin are almost the only ones which are distinguished by a smell and taste, sufficiently obvious to serve as characteristic marks. It has been noticed, that metals which are odorous, are also easily oxydated, in contact with moist air. Exposed to heat, many of the metals are melted, with more or less facility. Others, on the contrary, as platina, resist the most powerful furnaces, and are scarcely softened by the processes of metallurgy. There are several which melt above a red heat, such as tin, bismuth, and lead. These also crystallise, most readily;

and it is remarkable, that their crystals, whether met with in nature or obtained artificially, always have the same form.

We have seen, that metals are good conductors of caloric, and that they yield it to adjacent bodies, as readily as they receive it. They are likewise conductors of electricity; two insulated metals, as zinc and copper, which possess this property in an eminent degree, are placed, by contact, in two states of electricity, the zinc being positive, and the copper negative. On this circumstance, first observed by Volta, is based the apparatus which bears his name, viz. the *Voltaic* battery. Plates of zinc and copper, suitably arranged, compose the apparatus, by means of which the most astonishing decompositions and recompositions have been effected.

Common air, and even oxygen itself, when dry, exert no action on the metals. The results are different, if we bring to our aid, heat and moisture.

If we heat a metal, in a state of minute division, in contact with air, it absorbs oxygen gradually; at the same time, it loses its lustre and solidity, is changed into a brittle, dull substance, known, long ago, by the name of *calx*, but at present called *metallic oxyde*. This action is more obvious and speedy, in contact with oxygen gas, than with atmospheric air, of which oxygen constitutes only a fifth part. It is well known, that iron burns with energy in oxygen gas, with disengagement of heat and light, losing its lustre and consistence, and soon reduced to a brittle globule of protoxyde of that metal.

Exposed to common air, filled with moisture, the metals also experience a change, and are oxydated. Thus, on submitting iron to the twofold action of air and heat, we obtain an oxyde, which the ancient chemists called *astringent saffron of Mars*; and by exposing the same metal to moist air or the dew, an oxyde is prepared, that has very little resemblance to the last article, at least in point of oxydation, and it is called the *aperient saffron of Mars*.

Since it is necessary for the air to be moist, in order to act on metals, we are forced to admit, that water plays an important part in the oxydation of those bodies. The oxygen, consumed in the process of oxydation, is doubtless furnished, partly by the air and partly by the water suspended in the air, and rendering it moist. Water acts another part, also, viz. that of combining with the oxydes, in proportion as they are formed; and in proof of this position, it is known, that the oxydes which result from the action of moist air, are true hydrates. We remark further, that these oxydes almost always contain carbonic acid, which has evidently been furnished by the atmospheric air, in which it exists in very minute proportions.

Metals, for the most part, have such an affinity for oxygen, that they take it from the air, by the intervention of water or heat. There are several endowed with so powerful an affinity, that they decompose water, and separate oxygen from the hydrogen, sometimes without the aid of heat, as the metals of the alkalis; sometimes with the aid of a red heat, as iron, zinc, and manganese. There are some, on the contrary, which having but little affinity for oxygen, are reduced to the metallic state, by the action of a heat more powerful than that which sufficed to bring them to the state of oxydes; others have so feeble an affinity for oxygen, that they cannot be oxydated, but by means of acids, and are reduced by the slightest heat; such are silver and gold.

Respecting the classification of metals, there has been not a little diversity of opinion, among chemical writers; and each has chosen a method, according with some general, theoretical view of the science. I apprehend, that for all really useful purposes, we might examine these bodies in alphabetical order, and that in a course of lectures, no inconvenience would result to the lecturer from such an arrangement, while I feel equally sure, that the student would lose nothing. But as some regard to method seems to be called for in an elementary treatise, I shall pursue the order adopted by Mr Reid. The consideration of each metal shall be complete in itself; that is to say, all that is known, either in relation to the simple metallic body, or any of its compounds, will be brought into view in one and the same chapter. Thus, in treating of iron, we shall not only consider it as a simple body, but pursue it through all its connexions with other simple bodies, with acids, &c., before we enter upon the consideration of any other metal.

Pursuant to this arrangement, we shall divide metals into five orders, and each order into chapters, in conformity with the table of contents prefixed to this work.

Before I enter on the consideration of this subject in detail, it is proper to make a few remarks on the new metals, as they are called, with reference to the history of their discovery.

An examination of the works on chemistry, will show a constant uniformity in ascribing all the credit of discovering the alkaline and earthy metals, to Sir Humphrey Davy; and I am quite willing to do full justice to the memory of that eminent and highly useful philosopher. It does appear to me, however, that injustice has been done to Lavoisier in this matter, arising out of determined neglect or forgetfulness. My views of the science of chemistry were drawn from the anti-phlogistic system, as it is termed; and, as a matter of course, the writings of



the hero of that system have been peculiarly interesting to me. I have witnessed, therefore, with regret, the prevailing disposition to overlook the intrinsic merits of Lavoisier, and thus to class his *treatise* among the *obsolete* books of our science. I am very sure that many persons, who now talk about the discoveries of Sir H. Davy, in respect of the alkaline and earthy metals, have never read the important statements on the metallic nature of the alkalis and earths, which are to be found in that treatise; and it is high time that such persons had made themselves acquainted with those statements. It will not be denied, that a part of the book called the *Elements of Chemistry*, was not written by Lavoisier; but the whole is so entirely devoted to the anti-phlogistic chemistry, as to give the production a oneness of character, justifying the appellation so uniformly applied to it.

In a note, at the foot of page 213, the translator says, 'there are some experiments related in the Transactions of the Turin Academy, which give reason for suspecting, that soda is a modification of magnesia; this latter substance, according to the experiments detailed by Baron Born, seems to be a *metallic oxyde*. From analogy, we may presume, that *potash* is also a metallic substance, in some hitherto unknown state of combination. We shall thus exclude all the alkalis from the class of simple, elementary substances.' In page 217, Lavoisier thus writes: 'it is extremely probable, that barytes, which we have just arranged with the earths, is an oxyde; for in many experiments, it exhibits properties really approaching to those of metallic bodies. It is even possible, that all the substances we now call earths, may be only metallic oxydes, irreducible by any hitherto known process.' A section is added by the translator, on the *metallic nature of the earths*, in which, after detailing the processes by which the metallic bases of lime, barytes, and magnesia were repeatedly obtained, he uses the following language: 'should they be *eventually confirmed* by rigorous examination, a new light will be thrown on several of the most difficult parts of chemistry, in exact conformity with the predictions of Lavoisier, who supposes that those substances which have long been considered as primitive earths, are only metallic oxydes combined with oxygen, and that their reduction has hitherto been prevented, by the attraction which subsists between them and oxygen, being stronger than that between oxygen and carbon.' To this account, Baron Born adds, that 'he expects soon to learn, that the silicious and argillaceous earths are likewise metallic oxydes; and that in this case, the whole class of earths and stones will disappear from the mineral kingdom.'



The attentive reader, who has never, till the present time, consulted the neglected book to which I have referred, will be surprised to find the names *calcum*, *barytum*, &c. given to the simple metallic bases of lime and barytes. Yet such is the record of doings, more than thirty years ago.

There can scarcely be a doubt, that Professor Woodhouse (formerly of the University of Pennsylvania) was the first person who actually reduced potash to the metallic state. He was operating with pearlash and soot, exposed to a very high degree of heat in a covered crucible, and perceived, that water thrown on the mixture, when quite cold, caused it to take fire. He repeated the experiment with charcoal, and the same result followed. He had never seen potassium, nor heard of it; but who can now entertain a doubt, that he actually decomposed potash in the experiment which seemed to him quite unintelligible, and that the combustion which occurred under his notice, was the effect, so often seen in later times, of the contact of potassium with water.\*

All the facts and suggestions here detailed, were matters of record, in books accessible by all our modern authors, at least twenty years prior to the date of Sir H. Davy's alleged discoveries. I say *alleged* discoveries, for I am not willing to give all the honor to one man, when facts assure me that others are equally entitled to the meed of praise. Shall we be told, that the statements in Lavoisier's treatise, were regarded as mere speculations, or the fancied results of badly conducted experiments? Such a plea cannot be tolerated for a moment, and no one has ventured to make it. If the facts, as detailed, were doubted; if the smallest amount of hesitancy existed as to their authenticity, the proper course could not have been mistaken. But so far from any effort being made to invalidate these statements, not a work on chemistry that I have ever seen, even notices them; and hence it follows, that our young men, who have derived all their chemical information from recent authorities, do not know that the facts, whose merits I have been attempting to canvass, ever had existence.

Let it be conceded, for a moment, that the experiments referred to, were altogether delusive, and that their fallacy had been evinced, by frequent repetitions. All this would not nullify the claims of Lavoisier and his cotemporaries to our regard, touching the metallic nature of the alkalis and earths. No one can hesitate to believe, that the credit which has been awarded to Davy in these matters, had its origin in the sugges-

\* See Nicholson's Journal, vol. xxi. page 290.

tions accompanying the details of those experiments, and this fact ought to have been cheerfully acknowledged. But we have conceded, for the sake of argument, that which we had no right to grant; and until positive researches shall be made, to show the inaccuracy and fallacy of the experiments noticed in the *Elements* of Lavoisier, we are bound to say, that the credit of first reducing the earths and showing their metallic bases, belongs to men who flourished anterior to the performance of the brilliant galvanic experiments, that have given such lustre to the character of the English philosopher.

A word or two of explanation may be called for, in relation to the titles of the three first orders of metals. By *alkalifiable*\* metals, we mean, as the word imports, metals which are convertible into *alkalis*; by *alkaline geofiable*† metals, we mean such metals as give rise to the substances usually denominated *alkaline earths*; and by *geofiable* metals, those which produce the *earths*, strictly so called.

## ORDER FIRST.

### ALKALIFIABLE METALS.

#### CHAPTER I.

##### Of Potassium.

THE first individual who procured this substance, in its separate, metallic state, was Sir Humphrey Davy. In the year 1807, he exposed the hydrate of potash to the action of a powerful galvanic battery, first placing it between two discs of platina. In the experiment, oxygen appeared at the positive surface, and minute, metallic globules at the negative; and this result was found to have arisen from the decomposition of the potash into oxygen and potassium, whereby the alkali appeared to be a true metallic oxyde. A galvanic battery of two hundred pairs, is sufficient to show these globules, but we can obtain only a small quantity of the metal in this way.

\* Alkalifiable imports, the capacity of making alkalis.

† Geofiable means, the capacity of forming *earths*, and with the prefix *alkaline*, it denotes the power of making *alkaline earths*.

The process by which it is most easily procured, is by distillation from a mixture of potash or carbonate of potash and charcoal, without the addition of any iron filings or turnings. Brunner pointed out this process,\* but it has since been modified by Wohler; and the method of conducting it, with the form of apparatus which I shall now describe, appears to answer the end better than any other arrangement.

An iron pot is to be procured, made of the best malleable iron, of a circular form, about twelve inches long and six in diameter, and at least three-eighths of an inch in thickness. The top should be turned a little out, and a lid of the same metal and equal thickness, fitted accurately to it. This should be secured in its place, by an iron rod, passed through two holes in the upper part of the opposite sides of the pot. A hole at one side of the lid, admits a wide-bent gun-barrel, intended to convey the potassium to a receiver.

The mixture of charcoal and potash is prepared, most easily, by exposing six or seven pounds of cream of tartar (crude tartar may be used) to a red heat in large earthen or iron crucibles, till no more gas is disengaged, reducing it to powder, when cold, in a mortar. The gases that are disengaged, arise from the complete decomposition of the tartaric acid, its oxygen and hydrogen combining with part of its carbon and forming carbonic acid, carbonic oxide, and inflammable compounds of carbon and hydrogen; part of the carbonic acid remains in combination with the potash, with which the rest of the carbon is mixed. It is reduced to powder whenever it has become cold, and transferred immediately to the iron pot, that it may be prevented from attracting water from the air. Brunner states, that when the tartar is mixed with one-twelfth part of its weight of charcoal, a larger quantity of potassium is obtained.

The iron pot should be luted before it is used, as it is apt to be melted, or at least speedily destroyed by the action of the air, at the high temperature to which it must be exposed. The best method is to bind the iron vessel round on every side, in the first place, with iron wire; covering it afterwards to the depth of a quarter of an inch, with a stiff lute, made of good clay, with about a fifteenth part of iron filings and charcoal, using no more water than is necessary to allow it to be easily worked, and mixing all the materials with a little thread, (cut into small pieces not more than an inch long,) before adding the water. This is bound round again with iron wire, and the whole rubbed over

\* Quarterly Journal of Science, vols. xv. and xxii; or Bib. Aniv. xxi. and Annales des Mines, xii.

with an additional quantity of lute. The lower part of the apparatus should be luted first; it is easily covered with wire, by cutting a number of pieces a little longer than is necessary to allow them to go round both sides and the bottom of the iron pot, bending them in the form of the letter U, and putting a thick piece of the lute on the bottom of the pot, that they may be kept in their place till they have been fixed by the wire which is then wound round its sides. The mouth must be turned downwards all this time, and, when it has been inverted and the materials put in, the finger is drawn round the edge of the lid with a little common clay, (to prevent its adhering to the body of the iron vessel when it is at a high temperature,) and fixed in its place by the iron rod, which is passed above it through the rings that project from the top of the iron pot, a small piece of scrap iron (such as the end of a broken chisel) being hammered in between it and the lid, to drive it home. The bent gun-barrel (luted in the same manner as the iron pot, where it is to be exposed to the fire) is then introduced, and the ends of the iron wire, still projecting over the top of the pot, are then turned in over the lid and properly luted, taking care to protect the whole of the apparatus. If the furnace will not admit of the tube and the rest of the apparatus to be put in together, it will be necessary to place the iron vessel in the furnace, after the tube has been put in, and to finish the luting there. I have never seen this lute fail, when properly put on, though the apparatus may have been used immediately afterwards; it is better, however, to allow it to dry a day or two, previously, before the fire, filling up any cracks that may be observed, with a fresh quantity of lute.

After every thing has been properly adjusted, the fire may be put on. A little water comes away when the apparatus becomes red hot; soon after, carbonic oxyde gas is evolved, and when it is at a white heat, a very dense vapor is disengaged, which burns with a brilliant flame. The receiver, intended to condense the potassium, may then be fixed to the extremity of the gun-barrel, without the furnace, by a little plaster of Paris.

When the gas begins to be disengaged slowly, this arises in general from the tube being so obstructed, as to prevent it from passing out readily; the plug is then taken out of the tube, and the obstructing matter removed as completely as possible, but if the gas does not appear to increase in quantity, it will be better to withdraw the fire and allow the apparatus to cool. Too much caution cannot be taken in endeavoring to clear the tube, either during the distillation, or after the apparatus has been allowed to cool; for the tube being frequently obstructed while the ma-

materials are at a high temperature and still producing gas, it is obvious that a large quantity must be accumulated in a short time, and the moment the impediment to its free passage is removed, it often expands with explosive violence, and gives rise occasionally to serious accidents.

The operator ought always to be protected in the most complete manner, and never to be without a mask; processes like this should be excluded from class-rooms, where there is not an experienced assistant to give them that constant and undivided attention which they require.

Occasionally, too, when the tube becomes completely obstructed, the vapors of potassium are forced through between the lid and the top of the iron vessels, burning with a rich, violet-colored flame, easily distinguished, by its beautiful color, from the flame produced by the combustion of the fuel. The fire must be withdrawn, immediately, whenever this appears.

The calcined cream of tartar appears to be preferable to the mixture of carbonate of potash and charcoal, because it presents the potash in a state of most intimate mixture with the carbon of the decomposed tartaric acid. We are told, by Reid, that this mixture yields potassium at a lower temperature, than is necessary when iron filings are blended with charcoal.

If any one will compare the process of Brunner, as just detailed, with that of Woodhouse, as noticed above, he will perceive, at once, the propriety of our remark, touching the formation of potassium by the American professor. Woodhouse used pearl-ash (carbonate of potash) and soot, and exposed them to a high degree of heat in a crucible. A gas escaped, viz. carbonic oxyde, which arose from the decomposition of the carbonic acid, and the resulting new arrangement between the oxygen of that acid, its carbon, and a part of the carbon of the soot, whereby carbonic oxyde was formed. The heat being continued a little further and the apparatus then allowed to grow cold, it was found, that water thrown into the vessel excited actual combustion. The potash was, of course, decomposed by the deoxydating power of the carbon present, which, taking all its oxygen, left potassium in an uncombined state. All that the experiment of Woodhouse wanted, to complete its resemblance to that of Brunner, was the bent tube and receiver; the metallic product would have followed as a matter of course.

Gay Lussac and Thenard prepared potassium, by fusing the hydrate of potash and causing it to flow over iron turnings, heated to whiteness in a gun-barrel; the oxygen of the potash combined with the iron, and the potassium, being volatilised, passed over and was condensed in a receiver.



It is usually recommended to purify potassium, by distillation, in a common glass retort; but on repeating the operation, it will be found, that the temperature required for the volatilisation of the potassium, generally enables it to decompose the glass, which is often completely destroyed, and a considerable proportion of the potassium lost. A green glass retort is less apt to be destroyed, or any retorts which contain no lead. It will, perhaps, be better always to distil from a small iron vessel, or tube retort, heated by a chauffer.

In condensing potassium in the receiver, naphtha is the liquid that must be employed, as potassium can decompose every fluid which contains even a minute quantity of oxygen, combining with it and forming potash. The same fluid must also be employed to prevent the potassium from being oxydated by the air.

In the last edition of Turner's chemistry, the essential oil of copaiva is mentioned, as a substitute for naphtha. But Dr Gale, in a late number of Silliman's Journal, expresses his conviction, from actual experiment, that this article will not answer the purpose. I have used the *Seneka* oil, and find it quite as good as much of the drug sold for naphtha. If the latter fluid can be had, in the state of purity, no substitute need be sought for.

The equivalent of potassium is forty, and its specific gravity 0.865; at  $32^{\circ}$  it is hard and brittle, becomes soft and malleable at  $50^{\circ}$ , is fluid at  $150^{\circ}$ , and resembles mercury. At a red heat, it is volatilised.

The levity of potassium, at one time, constituted the chief ground of objection to its metallic character; but the more the subject has been examined, the less reasonable has this objection appeared, and no one, at present, ventures to bring it forward.

The leading character of potassium is, its strong attraction for oxygen. To illustrate this position, take a grain or two, and remove the naphtha adhering to it, by blotting paper, and place it, with a pair of pincers, on a piece of red hot iron. It will immediately take fire, combine with the oxygen of the air, and be converted into peroxyde of potassium. Exp.

Put another piece into a brass or copper cup, already at a high temperature, and introduce it, the moment it is inflamed, into a jar or bottle of oxygen gas, when it will burn more brilliantly than before. Throw a little water on the orange-colored matter that remains after either of these experiments; part of its oxygen is disengaged with effervescence, and potash remains in solution; on pouring a little of the infusion of red cabbage into the solution, it will immediately become green. Exp.

Throw a few pieces of potassium into a bason full of water; one portion of it immediately combines with the oxygen of part of the water, which is decomposed, forming potash, while the other uniting with the hydrogen, forms a very inflammable gas,

Exp. viz. potassiu-retted hydrogen, which takes fire as it is disengaged and is converted into potash and water. The potassium rolls along the surface of the water till the whole of it is oxydated, burning with a rich, rose-colored flame, and producing a very beautiful appearance.

Exp. Put another piece of potassium on a mass of ice; the same action takes place, and a similar light is produced.

Take a piece of potassium, (about a grain,) wrap it up in a small piece of paper and introduce it quickly into a glass-test, tube, inverted under water and full of this fluid. It will immediately rise to the top, and the moment the water reaches

Exp. it through the paper, part of it will be decomposed, the oxygen combining with it and forming potash, while an equivalent portion of hydrogen gas is found in the tube, and may be inflamed by applying a lighted match in the usual way.

### *Oxydes of Potassium.*

This metal combines with oxygen, in two proportions, and the compounds are called *protoxyde* and *peroxyde* of potassium. The first, or the *pure potash* of the shops, is always formed when the metal combines with oxygen at the ordinary temperature. If the metal be thrown into water, the potash is combined with that fluid, but if it absorb oxygen from the air, at the common temperature, the resulting compound will be *anhydrous*, or free of water. In forming the latter, the potassium must be cut into very small, thin slices, or the oxydation will be very tardy, and finally incomplete.

The protoxyde is a compound of oxygen 8, (one equivalent,) and 40 of potassium, (one equivalent). It is very soluble in water and in alcohol. The hydrate fuses before it is heated to redness. From the experiments of Gay Lussac and others, it appears that 40 grains of potassium decompose precisely nine grains of water, and that while one grain of hydrogen escapes from the latter in the gaseous form, the corresponding eight grains of oxygen combine with the metal. It was in this way, that the exact composition of the protoxyde was determined, and hence also was deduced its equivalent number.

The potash obtained when potassium is exposed to the action of air or water, as in any of the preceding experiments is very

pure; the expense of this process, however, must always preclude its being adopted for ordinary purposes.

The common process for obtaining potash, consists in preparing a solution of it, in the first place, from the carbonate, and evaporating it to dryness (to expel the water) in a silver or iron bason, carefully protected from the air, urging the heat afterwards till the dry mass is fused and begins to flow like oil. The process recommended by the different colleges, for the preparation of *potassa fusa*, is conducted in the same manner. For this purpose, pearlashes (the carbonate of potash that is met with in the shops) may be dissolved by rubbing them with four times their weight of water in an earthen mortar, and the solution decanted and mixed with a quantity of newly slaked lime, equal in weight to the pearlashes employed, boiling it for a few minutes and then filtering, to separate the solution of potash, that is procured in this manner, from the carbonate of lime with which it is mixed. The throat of the funnel is obstructed by a glass stopper or a piece of quartz, covering it closely with a piece of linen to prevent any lime passing through with the clear solution. As potash speedily attracts carbonic acid from the air and becomes carbonated, the solution must be protected from it, as much as possible, during the operation; the best method of effecting this is, by using a funnel with a narrow mouth which may be easily closed by a cork or stopper, and putting a small tube through the throat of the funnel, placing pieces of quartz or broken glass round it, and covering it with linen, so that while the solution of potash is dropping into the bottle below, air passes through the tube, at the same time, from the lower to the upper vessel, and supplies its place. This ingenious method was proposed by Dr Duncan, and is an excellent substitute for a more complicated apparatus. When a common funnel is employed, it should be covered with a plate or tin tray, and a towel thrown over the whole.

The whole of the solution of potash is not obtained in this manner, as a considerable portion of it adheres mechanically to the lime; to procure it, a small quantity of water is poured on the top of what remains in the funnel, and this causes the residue of the dissolved potash to pass slowly into the receiver below. This is continued, till a quantity of liquid is obtained, equal to five or six times the weight of the salt employed. It must be kept in glass bottles with good stopples.\*

In this process, the lime, having a greater affinity for carbonic acid than potash, immediately combines with it, carbonate of

\* This solution is the *liquor potassæ* of the Dispensatories.

lime being formed, which is insoluble and remains in the funnel, while the potash passes through in solution. It is impossible, in this manner, to remove the carbonic acid completely from all the potash; it is sufficiently pure, however, for most chemical purposes, when the precautions which have been mentioned are carefully attended to.

If the solution contains little carbonic acid, it will give a very slight precipitate with lime water, and scarcely present any appearance of effervescence with sulphuric acid. If it is free from lime, also, (which is occasionally present,) it will not be rendered turbid, on blowing through it with a bent tube.

It is from this solution that solid potash is usually formed by evaporating it to dryness, and then fusing the mass that remains, with a heat gradually increased, till it flows like oil, when it may be poured into moulds or on a plate of iron; it is then broken into pieces and put into a bottle, to prevent it from attracting water or carbonic acid from the air. In this state, it forms the *fused potash* or common caustic, of the different colleges. *Potash with lime* or the *milder caustic*, of the Edinburgh and Dublin colleges, is made by mixing the solution of potash (prepared from the carbonate) with slaked lime in fine powder, till it becomes of a thick consistence, after expelling about two-thirds of the water by evaporation.

When prepared in a silver vessel, it is usually of a white or greyish-white color, but when iron vessels are used, (which should always be bright and clean,) a portion of oxyde of iron is always mixed with it, giving it a brownish color, and separating from it, when it is dissolved in water. Hence the origin of the brownish matter which is usually deposited in dissolving this substance in water.

Though it is generally termed potash, or caustic potash, it is really a compound of dry potash and water, one equivalent of the alkali retaining one equivalent of water in combination, which cannot be separated by exposing it to a temperature far above a red heat. *Hydrate of potash* is, therefore, its strict chemical appellation.\* It has been found, also, that part of the potash attracts an additional quantity of oxygen during its fusion, and is converted into peroxyde of potassium, but this is not of any consequence, as the excess of oxygen is again disengaged, when it is dissolved in water.

Perfectly pure and dry potash has been prepared, as we remarked above, by exposing potassium, spread out into thin

\* It is also called the hydrated protoxyde of potassium, kali causticum, causticum commune acerrimum, lapis infernalis, &c.



layers, to the action of the air; but it has not been applied to any use. The common hydrate is what is used for almost all chemical purposes.

As the fused potash, prepared from the solution obtained by decomposing the carbonate, must always contain any carbonic acid that has been withdrawn by the lime, the latter must be separated by digesting it in strong alcohol, when this substance is required perfectly pure. The alcohol dissolves the potash, but leaves any carbonate that may be mixed with it; the spirit is afterwards separated by distillation.\*

Throw some pieces of fused potash into water; it will dissolve speedily, and considerable heat be produced, if the quantity of water is not great. The solution has a very acrid, Exp. alkaline taste, even when largely diluted with water.

Pour a little of the solution into a glass of water, tinged with the infusion of blue cabbage; it immediately becomes of a fine, green color, which speedily passes to a yellowish brown; Exp. the caustic potash, especially when the solution is strong, completely destroying the vegetable coloring matter.

Pour some of the solution of potash into two glasses of water, the one colored by an infusion of turmeric, and the other by a solution of litmus. The turmeric immediately becomes Exp. brown and the litmus remains unaltered; if, however, it should previously have acquired a reddish tint, it will be rendered blue.†

Fill several glasses nearly full of water, and drop into them small quantities of the solutions of a number of salts, as the muriate of barytes, the muriate of lime, the sulphate of magnesia, the sulphate of iron, the sulphate of copper, and the acetate of lead; then drop into each a small quantity of the solution of potash. It seizes the acid in combination with the different salifiable bases, all of which are completely separated, if a sufficient quantity of the alkali is employed, as potash has a much greater affinity for the sulphuric, muriatic, acetic, and most other acids,

\* This is usually termed, *pure potash by alcohol*.

† These experiments are illustrative of the leading features of *alkalis*, as noticed, when speaking of the vegetable alkalis. The peculiar, urinous taste, the reaction on vegetable colors, and the power of neutralising acids, seem to be the most distinctive marks of alkalis. Until recently, potash, soda, and ammonia, were spoken of, only as alkalis; the first two were called *fixed*, and the last, *volatile* alkali, because the former had a degree of permanency, not possessed by the latter. Potash was also called, at one time, the *vegetable* fixed alkali, because supposed to exist only in vegetables, while soda was called *mineral* alkali. These distinctions are, at present, obsolete, and never were of essential value.



when in solution, and at natural temperatures, and forms soluble salts with them; while the bases which it separates, are either insoluble or require a large quantity of water for their solution, and are thrown down in the form of precipitates.

The student should perform a number of similar experiments with the rest of the alkalis, &c., as he proceeds, till he becomes more familiar with the different substances with which they are incompatible, and their general chemical agency. Thus, potash is incompatible with acids, neutralising them immediately, and forming salts of various kinds, according to the nature of the acid; it is incompatible, in general, with soluble salts of the earths and common metallic oxydes, precipitating them from their solutions, and combining with the acids previously in combination with them; it decomposes the chloride and bi-chloride of mercury, (calomel and corrosive sublimate,) producing compounds which will be described under mercury, &c.; he need not attend, however, to the minutiae of its action on different bodies, till he come to study the different substances themselves.

Potash is used in medicine as a caustic,\* and is occasionally administered, internally, in calculous diseases. For chemical purposes, it is extensively employed in the arts, and is an important, analytical reagent. It is the basis of common soft soap, forms glass with silica, and its solution in water can dissolve a number of metallic oxydes.

The tests by which potash is most easily distinguished from soda, the substance that it is most likely to be mistaken for, are  
Exp. the chloride of platina and tartaric acid. The former gives a yellow-colored precipitate with salts of potash, when they are sufficiently concentrated, but does not affect those of soda; the action of tartaric acid has been already explained.

The potash of commerce is obtained, by digesting the ashes of burned wood in water, to extract the soluble matters which they contain, and boiling down the liquid afterwards, to expel the water. It consists, principally, of the hydrate of potash with carbonate of potash, and small quantities of the sulphate and muriate of potash, sulphuret of potassium, and earthy substances. When it is exposed to heat in a reverberatory furnace, stirring it constantly at the same time, it is more completely carbonated, and forms what is termed pearlash, from its

\* In the use of the common caustic of the shops, surgeons are in the habit of making it into a paste with water, and then rubbing it with some finely powdered quick-lime. This detaches the carbonic acid from the potash, and renders it sufficiently pure.

acquiring a white color. A small quantity of alkaline matter, sufficient at least to indicate its presence, may be obtained, in solution, by pouring water on the ashes of a wood fire and filtering the liquid, which will then turn the infusion of red cabbage to green.

It has happened, that strong solutions of potash and the other alkalis have been swallowed by mistake; and unless the accident be speedily remedied, death must follow, if the solutions be strong. The best mode of treating these cases is to give, freely, the oil of almonds, or sweet oil, which tends to change the alkali into a soap. Weak acids have been advised, but the oils are preferable.

The *peroxyde of potassium* is composed of oxygen 24, (three equivalents,) and potassium 40, (one equivalent). It is made, by burning potassium in the open air, or in oxygen gas. The effect of caloric is to divide the particles of the metal, and so present a larger surface for the oxygen to act upon. The product is an orange-colored substance, which is decomposed, readily, by water; its oxygen instantly escapes, and the metal takes a sufficient quantity of oxygen from the water, to form the protoxyde of potassium, or potash.

A more easy method of forming this peroxyde, was ascertained, in Philadelphia by Dr Bridges, and in London by Mr Phillips, in the year 1827. Nitrate of potash is exposed to a red heat, as long as any gas comes over. When water is poured on the residuum, oxygen gas escapes, and we are thus furnished with another mode for obtaining this gas. After the gas has escaped, hydrate of potash remains in the vessel. Now the potash in the nitrate, is in the state of protoxyde, but as the nitric acid is decomposed by the red heat, an abundant supply of the oxygen is afforded, to convert this protoxyde into the peroxyde.

This compound has not been applied to any useful purpose, in the practice of medicine or the arts.

Having thus given a brief account of the combinations of potassium with oxygen, I proceed to notice the products arising from its union with other simple, non-metallic bodies, and with the acids.

Potassium combines with hydrogen, forming *potassiuretted hydrogen gas*. It is produced, when hydrate of potash is decomposed by iron at a white heat, and also, when potassium is burnt on the surface of water. This gas inflames, spontaneously, in air or oxygen gas, its hydrogen being changed to water, and its potassium into potash. If it be kept for some hours over mercury, nearly all of its potassium is deposited.

Another compound of potassium and hydrogen, was made by Gay Lussac and Thenard, by heating the metal in hydrogen gas. It is a gray, solid substance, easily decomposed by heat or contact with water, and not inflamed in oxygen gas. It is called *hydroguret of potassium*.

*Sulphuret of Potassium.* This may be obtained, by putting a small quantity of potassium into a glass tube, covering it with a little sulphur, and exposing it to heat; the sulphur and potassium combine, and heat and light are evolved during the combination. Sulphuret of potassium is usually prepared, by decomposing the salt called sulphate of potash with charcoal; the sulphate is reduced to an impalpable powder, mixed with a fourth of its weight of charcoal, also in fine powder, and the mixture exposed to a strong furnace heat. There is disengaged a quantity of carbonic oxyde and carbonic acid gases, the whole of the oxygen of the sulphuric acid and potash combining with the carbon, while the sulphur of the acid and the potassium of the potash, join to form sulphuret of potassium.

The annexed diagram will illustrate the process, supposing no carbonic acid to be formed, which is probably the case when carbon is present, in sufficient quantity.

Before decomposition.		After decomposition.	
24 or 4 equiv. carbon	{ carbon	6	14 carbonic oxyde.
	{ carbon	6	
	{ carbon	6	
	{ carbon	6	
40 or 1 equiv. sulphuric ac.	{ oxygen	8	14 carbonic oxyde.
	{ oxygen	8	
	{ oxygen	8	
	{ sulphur	16	
48 or 1 equiv. potash	{ oxygen	8	56 } or 1 equiv. sulphu- ret of potassium.
	{ potassium	40	
112		112	112

Sulphuret of potassium may also be prepared, by heating the carbonate of potash with sulphur, in a large crucible; the heat ought to be applied cautiously, and continued till the mixture is completely fused. It is thus, that the different colleges prepare what is commonly called *sulphuret of potash*; but it is in reality, a mixture of sulphuret of potassium and other compounds, produced by the action of the sulphur on the potash and its oxygen. As in color it resembles the liver of animals, it is called *hepar sulphuris*.

This sulphuret is composed of sulphur 16, (one equivalent,) and 40 potassium, (one equivalent). Berzelius speaks of several other sulphurets of potassium, but they are not of sufficient importance to engage our attention.

The sulphuret now before us, is called the protosulphuret, and when digested in water, gives rise to double decomposition, and the resulting compound is the hydrosulphuret of potash, or potash combined with sulphuretted hydrogen, as the following diagram evinces.

Before decomposition.		After decomposition.	
9 or 1 equivalent of water	{hydrogen	1	17 sulph. hyd.
	{oxygen	8	
56 or 1 equiv. of sulph.	{sulphur	16	48 prot. potas.
of potassium	{potassium	40	
—	—	—	—
65	65	65	{hydrosul.
—	—	—	{of potash.

The forty-eight parts or one equivalent of protoxyde of potassium, or potash, unite with the seventeen parts or one equivalent of sulphuretted hydrogen; and the compound is called *hydrosulphuret*, because it is important to designate both the simple substances with which the potash is combined.

But besides the hydrosulphuret of potash, we have the *sulphuretted* hydrosulphuret, by which is meant, the hydrosulphuret combined with one equivalent of sulphur. This is exemplified by the following diagram, showing the action of 64 grains of sulphur on 114 grains of the hydrate of potash, boiled in water. The whole of the sulphur and potash is dissolved, and the solution is of a dark greenish color. The diagram leaves out of view one equivalent of water, belonging to the hydrate of potash.

9 or 1 equiv. of water	{oxygen	8	88 {hyposulphite of potash.
	{hydrogen	1	
64 or 4 equiv. of sulphur	{sulphur	16	
	{sulphur	16	
	{sulphur	16	81 {sulphuretted hydrosul. of potash.
	{sulphur	16	
96 or 2 equiv. of potash	{potash	48	
	{potash	48	
—	—	—	—
169	169	169	
—	—	—	

In this process, two equivalents of sulphur unite with one of oxygen, to form hyposulphurous acid, which joins one equivalent



of potash, making the hyposulphite of potash. One equivalent of sulphur and one of hydrogen forms the sulphuretted hydrogen, which, joined to the other equivalent of potash, forms the hydrosulphuret of potash. This latter, combined with the remaining equivalent of sulphur, gives rise to the sulphuretted hydrosulphuret of potash.

Another equivalent of sulphur might be dissolved in the same quantities of water and alkali. The solutions have a deep, amber color when they become clear, attract oxygen from the air, and are decomposed by the acids, with disengagement of sulphuretted hydrogen, and a free precipitation of sulphur.

The common sulphuret of potash is a hard, brittle mass, with an acrid and bitter taste. When dry, it has no odor, but if moistened, it yields the well known smell of sulphuretted hydrogen. As acids, in common with acidulous and metallic salts, decompose it, they are to be regarded as incompatibles. It has been used in medicine, internally and externally, and from its chemical action on metallic salts, it has been employed as an antidote to those that are poisonous.

*Phosphuret of potassium* may be formed by the action of potassium on phosphorus, with the aid of a moderate heat. Water converts it into potash and perphosphuretted hydrogen gas, which inflames at the moment of its formation.

*Chloride of Potassium.* If this metal be placed in an atmosphere of chlorine, it takes fire spontaneously, and burns with more brilliancy than in oxygen gas, and a chloride is the result. The same compound is generated, when potassium is heated in muriatic acid gas, and hydrogen is evolved. The acid gas is simply decomposed, its chlorine joining the metal and its hydrogen escaping. This chloride is the residuum of the decomposition of chlorate of potash by heat, as we showed when treating of oxygen gas.

Chloride of potassium has a saline and slightly bitter taste. It requires three parts of water, at  $60^{\circ}$ , for solution, but is rather more soluble in hot water. It is composed of chlorine 36, (one equivalent,) and potassium 40, (one equivalent). *Iodide of potassium* will be noticed under the article *hydriodate of potash*.

*Cyanide of Potassium.* This is a compound of cyanogen and potassium, which M. Chevallier directs to be prepared by calcining the prussiate of potash, (ferrohydrocyanate, to be treated of hereafter,) and then separating the cyanide from the carburet of iron, by means of alcohol. Distillation then yields this compound, free of alcohol, very pure and white.

The *salts* of potash next come under review; but before we



enter upon their examination, it is proper to premise a few remarks on the nature of salts, in general.

The term *salt* was, for a time, limited to bodies having a strictly saline taste. But we now apply it, without reference to taste, color, or smell. Chemists, for the most part, mean by this term, a definite compound of an acid, and an alkaline or salifiable base. By a *neutral* salt, is meant, a compound in which neither the acid nor base is in excess; the peculiar properties of the acid and base being alike destroyed.

A *deliquescent* salt is one which readily absorbs moisture from the air, and becomes more or less liquid. An *efflorescent* salt is one that gives out its water to the surrounding air, and is reduced to a pulverulent form. The water thus lost is called the water of crystallisation.

If the solution of a salt redden litmus paper, we say, that the acid is in excess; if it brown turmeric paper, we infer that the alkaline base is in excess. But the relative proportion of acid in a salt may be greater than that of the base, and yet the solution of the salt shall exhibit alkaline properties; this is the case with the well known salt, borax. The term *triple* has been applied to salts having two bases joined to one acid, or two acids united to one base. Berzelius employs the word *double*, instead of triple, but I prefer the latter, as rendering innovation needless, and, at the same time, serving all useful purposes.

Salts usually exhibit specific, crystalline forms, but as fatal mistakes might arise from attempting to rely upon distinctions drawn from this source, and as crystallography is peculiarly within the province of mineralogy, I have not deemed it important to devote any portion of this work to the consideration of that subject. I feel confident, that, for all practical purposes, better modes of distinguishing between saline bodies are within our reach, than the interesting study of crystallisation can afford.

The first salt of potash that claims our attention is the *nitrate of potash*. As its name imports, this salt is a compound of nitric acid and potash. Its equivalent is 102, being composed of nitric acid 54, (one equivalent,) and potash 48, (one equivalent). It is soluble in seven parts of water at 60°, and in its own weight at 212°; its solubility is further increased, by adding a small portion of muriate of soda (common salt) to the water. In all cases of its solution, there is a considerable reduction of temperature. Alcohol does not dissolve it. It is well known by the names of *nitre* and *saltpetre*.

We may show the formation of this salt, by adding nitric acid to a solution of common potash, which is always combined with more or less carbonic acid, in form of a carbonate or subcarbo-

nate. The nitric acid seizes the potash and displaces the carbonic acid, which passes off in the act of effervescence. The acid should, of course, be added until the potash is entirely neutralised, and then the solution should be evaporated till its bulk is reduced to about four times the amount of nitric acid employed; on cooling, crystals of nitrate of potash will be formed.

But the great quantities of nitre manufactured for various uses in the arts, medicine, &c. are not thus prepared. In some parts of the world, especially in the East Indies, nitre is an abundant natural product, appearing on the earth's surface, in an efflorescent form. In the same manner, it is found in Spain, Naples, and elsewhere, in considerable quantities. But in this country, it is seldom seen under those circumstances, although large quantities of what is called *calcareous nitre*, are met with in several caves and other locations, in Georgia, Kentucky, &c. More commonly, it is formed in *artificial nitre beds*, by mixing lime and its carbonate with animal and vegetable substances, and exposing them to the air. These perishable matters, by decomposition, yield nitrogen and oxygen, which readily unite, in their nascent form, and constitute nitric acid, the air also contributing a supply of the same elementary matters. The acid thus generated, takes up the lime, and gives rise to a salt of easy solution in water, viz. the nitrate of lime, which is precisely the same, in its chemical character, with the calcareous nitre just referred to. This process is conducted in composts and similar heaps, in farm-yards and other places. As the nitrate of lime is thus formed, and as it is easily dissolved in water, it is not difficult to decompose it, and obtain the nitrate of potash. The mass may be placed in hogsheads furnished with cocks, and water being poured in, carries the nitrate of lime into other vessels, where it is mixed with a strong potash ley, or a solution of carbonate of potash. The process of double decomposition instantly commences, and the result is, a precipitation of insoluble carbonate of lime, and the formation of nitrate of potash, which is held in solution; evaporation and crystallisation yield the salt in abundance.

Under every circumstance in which nitrate of potash is formed, it is liable to mixture with more or less common salt, (muriate of soda,) excepting when made by the direct union of nitric acid and potash. The presence of this foreign matter may be ascertained, by adding to the solution, a few drops of a solution of nitrate of silver; if muriate of soda be present, there will be a white, curdy precipitate, consisting of muriate or chloride of silver.

As nitrate of potash is of extensive use in medical practice,

it is important to know the articles with which it is incompatible. Among these, we mention *alum*, *sulphate of magnesia*, *sulphuric acid*, *sulphates of zinc*, *copper*, and *iron*, because, when their solutions are mixed with solutions of nitre, decomposition takes place. It might be supposed, that sulphate of soda would also be incompatible, but as no decomposition occurs at ordinary temperatures, and even at  $32^{\circ}$  only in a partial degree, it may be combined with nitrate of potash with propriety, if the mixture be desirable.

When the refrigerant action of nitre is desired, the solution should be given immediately; otherwise, on the principles of the equilibrium of heat, the effect will be lost.

The taste of nitre, though saline, is accompanied with an impression of coolness, which might serve to distinguish it from the common Glauber's salts, (sulphate of soda). Not a few mistakes have occurred, from the want of tact to distinguish between these salts; and I knew a graduate in medicine who took an ounce of nitre, in mistake for Glauber's salts, the day after he received his diploma, and his life was painfully jeopardised by the blunder. Violent vomiting, bloody discharges, and cramp were induced, from the effects of which he was happily saved by the action of opiates, aromatics, and fomentations, liberally administered. It will not be doubted, that the sufferer, however deficient before, was not slow to learn the distinctive marks of these salts.

The chemical characteristics of nitre, that may be shown very easily, are, its power of animating the combustion of fuel, its deflagration when thrown on burning coals, the evolution of suffocating nitrous vapors, when the salt is heated in contact with strong sulphuric acid, and the precipitation, of a yellow color, on adding chloride of platinum to its solution. Not one of these features belongs to Glauber's salts; and when in addition to these characteristics, the different impressions produced by tasting the two salts are remembered, apart from their crystalline characters, there can be no difficulty in making the proper distinction.

In addition to the employment of nitre in medical practice, its uses are extensive in the arts, in chemistry, metallurgy, &c. It is employed in the art of dyeing, in manufacturing gunpowder, and in curing meat; its action in the latter case is not well understood, but most probably depends much on the generation of cold. It is also employed in the formation of nitric and sulphuric acids, and a number of fulminating and deflagrating compounds. It enters into the composition of fluxes for the reducing of metals, is used for oxydating a number of sub-

stances, and occasionally, for making freezing mixtures. Of its use in the preparation of oxygen gas, we have already spoken.

A few experiments may be made with it, to exhibit its general properties.

If we mix 300 grains of nitre, 40 of charcoal, and 60 of sulphur, all in fine powder, we shall have the exact proportions of some of the best kinds of shooting powder; and although the use of suitable machinery is requisite, in order to give the mixture the proper appearance, it will be found to deflagrate on applying a lighted taper.

The *portfire* used for artillery is made of three parts of nitre, two of sulphur, and one of gunpowder, well mixed and rammed in cases; a small quantity will suffice to show the appearance which it presents on burning.

*Signal lights* are generally composed of nitre and sulphur, with a small quantity of some metallic sulphuret, as that of arsenic or antimony. Mix 600 grains of nitre with 200 of sulphur and 100 of yellow sulphuret of arsenic; put the mixture into a cone of paper, and touch it (out of doors or under a large chimney or hood) with a red hot iron; it deflagrates rapidly, with a brilliant, white light.

Mix 100 or 200 grains of the sulphuret of antimony, with the same proportions of sulphur and nitre, and inflame it in the same manner. It deflagrates immediately, with a vivid light, which has a bluish tinge.

Mix eight grains of powdered nitre with four grains of phosphorus, cut into pieces; place the mixture on an anvil, and strike it smartly with a hot hammer. A violent detonation will be the result.

Mix ten grains of nitre with three grains of charcoal, both in fine powder, and throw the mixture on a red hot shovel. A very brilliant combustion, attended with a loud explosion, will be the consequence.

If nitrate of potash be melted with a gentle heat, and poured into circular, flat moulds, of a half inch in depth, it forms the article sold in the shops as *sal prunelle* or *crystal mineral*. The only change effected in the salt, in this operation, is the removal of its water of crystallisation.

If equal weights of nitre and sulphur be put into a red hot crucible, the compound called *sulphas potassæ cum sulphure*, is formed. The sulphur reacts on the oxygen of the nitre, in a way that is not well understood. It is certain, that the product is different from a mere mixture of sulphate of potash and sulphur, although its name indicates that state of



union. Paris says, it contains sulphate, bi-sulphate, and sulphuret of potash. It is known, also, by the name of *sal polychrest* or *salt of many virtues*.

The *nitrite of potash* is composed of nitrous acid and potash. If nitre be exposed to a red heat, so as not to effect entire decomposition, its nitric acid will lose so much oxygen, as to be changed into nitrous acid. This salt is of no importance, so far as we know.

*Sulphate of Potash.* This article has been known by the names of *vitriolated kali*, *vitriolated tartar*, *sal de duobus*, &c. It is composed of sulphuric acid 40, (one equivalent,) and potash 48, (one equivalent,) and is soluble in sixteen parts of cold and five of boiling water, but insoluble in alcohol. The crystals are slightly efflorescent, and when heated, they decrepitate.

This salt is obtained from the residuum of the distillation of nitric acid, which we stated to be, a bi-sulphate of potash, or potash combined with two equivalents of acid. This residuum is soluble in water, and if common potash be added to neutralise the excess of acid, and a part of the water be expelled by evaporation, it will yield small crystals of the sulphate, on cooling. Carbonate of lime may be employed, instead of carbonate of potash, to remove excess of acid; and for this purpose it should be added until effervescence ceases. The following diagram shows the action that takes place, supposing all the potash to be in the form of bi-sulphate.

Before decomposition.			After decomposition.	
128 or 1 equiv. of bi-sulph. pot.	{	potash 48 sulph. ac. 40 sulph. ac. 40	88 sulph. potash.	
50 or 1 equiv. of carb. lime	{	carb. ac. 22 lime 28	22 carb. acid.	68 sulph. lime.
<hr/> 178		<hr/> 178	<hr/> 178	<hr/> 178

As the sulphate of lime is partly soluble in water, and the sulphate of potash still more so, some difficulty occurs in attempting to separate them. If we use carbonate of potash instead of carbonate of lime, we shall have no other product than the neutral sulphate of potash. This is shown by the following diagram.



Before decomposition.

After decomposition.

128 bi-sulphate of potash	{	potash	48	—	88 sulph. potash.
		sulph. ac.	40		
70 carb. potash.	{	sulph. ac.	40	—	88 sulph. potash.
		carb. ac.	22		22 carb. acid.
		potash	48		
198			198		198

The nitric and muriatic acids partially decompose the sulphate of potash, and convert a portion of it into bi-sulphate; of course, they are incompatible. So also are lime and its compounds, corrosive sublimate, nitrate of silver, acetate and subacetate of lead.

Formerly, this article was used in the practice of medicine, in combination with rhubarb; but it is not employed in this country, excepting in the well known Dover's powder, where it seems to have been added, simply on account of its triturating and dividing agency, when rubbed with opium and ipecacuanha.

*Bi-sulphate of potash* or the supersulphate, is the residuary mass found in the retort after making nitric acid, by decomposition of saltpetre. It is known in commerce by the title of *sal enixum*. It is composed of two equivalents of sulphuric acid, 80, and one equivalent of potash, 48. It is soluble in twice its weight of water, as well as in alcohol. It has been used as a cathartic, in union with rhubarb.

*Carbonate of Potash.* The compounds of carbonic acid and potash, have been named without sufficient regard to their chemical constitution. Hence, the name *subcarbonate* is often used, when the prefix *sub* should be dropt. We shall notice two compounds, called *carbonate* and *bi-carbonate*. A striking difference between these, is, that the former is obtained in the form of fine grains, which absorb moisture rapidly from the air, and run into a dense, oily solution, that was formerly called *oleum tartari per deliquum*; while the latter is found in regular crystals, that are permanent in the air.

The *carbonate* of potash has been called *subcarbonate*, because it has alkaline properties, although the acids are joined in equal proportions; that is to say, the salt contains twenty-two of carbonic acid (one equivalent) and forty-eight of potash (one equivalent). The atomic constitution, which now very properly regulates our nomenclature, requires us to prefer the appellation which stands at the head of this article. The old books call it *kali preparatum*, *sal absynthii*, *sal tartari*, &c.; and I have seen an apothecary shop, in which were bottles labelled with several

of these names, as though there were really any difference in the contents of the bottles, such as the variety of labels seemed to indicate.

The impure carbonate, as obtained from a variety of vegetables, contains some silicious matter, sulphate and muriate of potash, alumina, and other articles, but they are present in too small quantities to prejudice the alkaline action of the carbonate.

The *pearlash* and *potash* of commerce are impure carbonates; the former named from its pearl-white color, and the latter, from the fact that the process is carried on in iron *pots*. We may make a solution sufficiently pure for ordinary purposes, by dissolving pearlash in water; the pure carbonate will be held in solution, while all the adventitious matters, being less soluble in cold water, remain undissolved. The pure carbonate requires twice its weight of cold water for solution, but is insoluble in alcohol. It unites with oils and forms soaps.

A pure carbonate may be made by decomposing the cream of tartar (bi-tartrate of potash) by means of heat. The tartaric acid, in one equivalent of cream of tartar, contains forty-eight parts of carbon, eighty of oxygen and four of hydrogen; and as the carbonic acid, in one equivalent of carbonate of potash, contains only six of carbon and sixteen of oxygen, it is easy to perceive how a carbonate of potash may be produced by the decomposition of the bi-tartrate. All the excess of carbon, oxygen, and hydrogen is carried off, in form of gas or vapor, by the heat employed in the process.

The carbonate of potash must be kept in close vessels, to prevent the absorption of moisture. A solution of this salt changes the infusion of blue cabbage to green, and as it is much less caustic than pure potash, the green color remains for a considerable time.

The carbonate of potash is sometimes adulterated with lime, which is easily detected by oxalic acid. The purity or strength of the salt may be estimated, by the quantity of nitric acid of a given density, requisite for the saturation of a given weight.

The incompatibles of carbonate of potash are, all *acids* and *acidulous salts*, *borax*, *muriate of ammonia*, *acetate of ammonia*, *alum*, *sulphate of magnesia*, *lime water*, *nitrate of silver*, *ammoniu-rect of copper*, *muriate of iron*, *calomel* and *corrosive sublimate*, *acetate of lead*, *tartar emetic*, *tartarised iron*, *sulphates of zinc*, *copper*, *iron*, &c.

The carbonate of potash is largely used in the practice of medicine, to form saline and effervescing draughts. A simple, saline draught, made of carbonate of potash, is a watery solution

of the salt, and it may become effervescent, by the decomposing power of acid in the stomach. But the effervescing mixtures are usually made, by adding to a watery solution of the carbonate, a portion of lemon juice or other acid. The mixture is followed with brisk effervescence, and is taken in that state.

*Bi-carbonate of potash* is a compound of 44 carbonic acid (two equivalents) and 48 potash (one equivalent). It has been called *supercarbonate*, because the acid is in excess, although its solutions have decidedly alkaline properties. It has also been called a carbonate, by those who regarded the other compound as a subsalt.

The bi-carbonate is obtained in regular crystals, which are very soluble in four parts of cold, and in five-sixths of their weight of boiling water; but they are insoluble in alcohol. When the crystals are dissolved in boiling water, they are decomposed, and converted into carbonate of potash; the same result follows the exposure of the bi-salt to a red heat, and this is a certain mode for procuring the carbonate perfectly pure and dry.

The bi-carbonate may be made in a Woulfe's or Nooth's apparatus, by passing a stream of carbonic acid through a solution of the carbonate. A very gentle heat, not sufficient to expel the excess of acid,\* is required to evaporate the solution, after which it crystallises slowly, in prisms. One hundred and one grains of crystallised bi-carbonate, are decomposed by seventy-five grains of crystallised tartaric acid and seventy-six grains of crystallised citric acid. Carbonic acid rapidly escapes during this action, while tartrate or citrate of potash remains in solution. Thirty or forty grains of tartaric acid and a corresponding quantity of the bi-carbonate, dissolved in separate glasses of water, together with a few pieces of sugar, form an agreeable and refreshing draught, when mixed together.

By reason of the excess of carbonic acid in this salt, it is much better suited to the condition of the stomach, than the carbonate, or any other alkaline combination. Its incompatibles are the same that have been named under the article, carbonate of potash.

In addition to the carbonate and bi-carbonate, Dr Thomson has described a *sesqui-carbonate*, composed of one equivalent of potash, one equivalent and a half of carbonic acid, and six equivalents of water. It was discovered by Dr Nimmo of Glasgow, but is not likely to be of any practical importance. It has been obtained by subjecting the bi-carbonate, in solution, to a degree of heat insufficient to reduce it to the state of carbonate.

\* A boiling heat drives off all the excess of acid.

*Acetate of Potash.* This is a compound of acetic acid and potash, and is one of the articles from which that acid is obtained by decomposition. It is also known by the names of *acetas kali*, *sal diureticus*, *terra foliata tartari*, *sal sennertii*, *magisterium purgans tartari*, *sal essentielle vini*, *sal digestivus sylvii*, &c.

It may be made, by adding pure acetic acid to the carbonate of potash in solution, until the mixture is perfectly neutral; after this, the water is to be expelled by evaporation. It is generally seen in masses of a foliated texture, and it is extremely deliquescent. If the salt be fused, it becomes a solid, crystalline mass, on cooling. At a high temperature, it is entirely decomposed. It has very little odor, its color is white, when pure, and its taste is sharp and pungent. One ounce of pure water, at 60°, dissolves 404 grains, and the solution soon undergoes spontaneous decomposition. It is soluble in alcohol, in the proportion of four parts of that fluid to one of the salt.

The composition of acetate of potash is, acetic acid 51, (or one equivalent,) and potash 48, (one equivalent).

This salt is decomposed by *tamarinds*, and most of the *subacid fruits*, by nearly all the *acids*, as well as almost every variety of *neutral salt*, whether *alkaline*, *acid*, or *metallic*; of course, all these substances are to be regarded as incompatibles.

The acetate of potash is decomposed by the digestive organs, owing to the presence of acid matters. One of its earliest names, viz. *sal diureticus*, points to its peculiar action, in the hands of practitioners, in former times. In addition to its diuretic action, it is also cathartic, but both these properties are, doubtless, dependent on the decomposition effected in the stomach.

Acetate of potash is sometimes adulterated with tartrate of potash, (soluble tartar); but as the bi-tartrate is insoluble, we have only to add tartaric acid in solution, in order to expose the fraud. The brown tinge, sometimes seen in this salt, is owing to some extractive matter, or slight traces of copper. There are two compounds of potash with tartaric acid that are used in medicine, and admit of other important applications. They are called the *tartrate* and *bi-tartrate*.

*Of the Bi-tartrate of Potash.* This salt is composed of 132 tartaric acid (two equivalents) and 48 potash (one equivalent). It is obtained in form of irregular, brittle, crystalline lumps, which, when reduced to powder, are termed *cream of tartar*. Its taste is harsh and acid, and its solutions redden litmus paper. It is called, also, *supertartrate of potash*, because the acid is in excess.

It is prepared, for the most part, from the crude tartar that is

deposited on the sides of wine casks, and which consists of the bi-tartrate and the coloring matter of the wine. This tartar or argol is held in solution in the wine, when first put into the casks, but, by long standing, it seems to purge itself of this acidulous matter, in order to acquire a better quality. The elementary principles of the tartaric acid unquestionably existed in the grape or other juice, employed in manufacturing the wine, and potash is well known to be a vegetable product. The alcohol, formed by the vinous fermentation, is not able to hold but a small portion of the salt in solution, and hence its tendency to precipitation.

The mode of separating the bi-tartrate from crude tartar, is as follows. The crude mass, being reduced to coarse powder, is boiled in four times its weight of pure water, after which albumen and wood ashes are added. The former coagulates and collects various impurities, whilst the latter occasions an effervescence, by which these matters are raised to the surface, whence they are removed by frequent skimming. Sometimes aluminous earth is used to precipitate the coloring matter, and sometimes the color is destroyed by well burnt charcoal.

From the solution, thus purified of foreign matters, the bi-tartrate may be obtained by evaporation and crystallisation. It may also be made by adding tartaric acid to a solution of tartrate of potash, (soluble tartar,) until an insoluble matter is thrown down, which is the bi-tartrate.

This salt requires 125 parts at  $60^{\circ}$ , and 30 at  $212^{\circ}$ , for solution, and hence it is usually called an insoluble salt. It is slightly soluble in alcohol.

The watery solution was long ago observed, by Berthollet, to undergo spontaneous decomposition, after being kept for some time; a mucous matter is deposited, and there remains a solution of carbonate of potash, colored with a little oleaginous matter.

Vogel ascertained that the bi-tartrate might be rendered soluble by combination with boracic acid, and, accordingly, he introduced a formula entitled, *tartras acidulus potassæ solubilis, admixto acido boracico*. It has been since called, the *borotartrate* of potash, and may be made by heating a mixture of 30 parts boracic acid, and adding to it, gradually, 120 parts of bi-tartrate, taking care to shake the whole frequently. By continuing the heat, the whole will be liquefied, and ultimately a pulverulent mass will be formed.

But as the diuretic action of the bi-tartrate is the most desirable effect to be had by its use, and as this depends on its comparative insolubility, it ceases to operate thus, so soon as it acquires a soluble form. As a neutral salt, it will act upon the



bowels, and be carried out of the system before it can have time to affect the urinary secretions. It is, therefore, undesirable to adopt the formula of Vogel, or any other by which the appropriate operation of the bi-tartrate might be changed.

All the *alkalis* and *alkaline earths*, and the *mineral acids*, are incompatible with this salt. As it decomposes the bi-carbonates of potash and soda, it is sometimes added to these, in solution, to form effervescing, purgative draughts. It is used in the formation of the tartrate of iron, emetic tartar, and Rochelle salt.

The only substance with which the bi-tartrate is adulterated, is the bi-sulphate. It may be detected by its superior solubility, and also by the addition of barytic water.

The bi-tartrate is used to form the *white* and *black fluxes*. When equal parts, by weight, of nitre and cream of tartar are mixed together and thrown into a red hot crucible, a rapid deflagration ensues, and a white matter remains, which is usually called *white flux*. It is a pure carbonate, the nitric and tartaric acids of the two salts being completely decomposed, and part of the carbonic acid, thus formed, remaining in combination with the potash.

The *black flux* is formed from the same materials, presented in different proportions. If two parts of bi-tartrate are mixed with one of nitre, and the mixture thrown quickly, in successive portions, into a red hot crucible, furnished with a lid which should be closed, leaving only a small aperture for the escape of gas, a black mass will remain in the crucible, which is the black flux. It is composed of carbonate of potash and charcoal, the oxygen of the nitric acid (of the nitre) not being sufficient to carry off the excess of carbon contained in the tartaric acid.

The white flux is used in several chemical operations, to impart fluidity to different substances with which it is mixed at a high temperature; the black flux may be employed for similar purposes, but it is used, chiefly, as a deoxydating agent, as in the reduction of arsenious acid. When a pure solution of carbonate of potash is required as a test, it should always be prepared from cream of tartar. In forming both these fluxes, a crucible should be selected that will hold twice the quantity of materials that may be required.

The salt, formed by saturating tartaric acid with potash, is a tartrate, and is called the *neutral tartrate of potash*, and on account of its easy solubility, it is more commonly styled, *soluble tartar*. We seldom see this salt in any thing like a crystalline form, but meet with it, usually, in the state of grains or powder. It is composed of one equivalent of tartaric acid, 66, and one of potash, 48. It is quickly dissolved, in three or four parts of water, and its solution, thus made, undergoes spontaneous de-

composition, and a carbonate of potash is the result. It is also quite soluble in alcohol.

The neutral tartrate is made, by adding a solution of potash to a solution of cream of tartar, until the excess of acid is neutralised; or it may be made, by adding powdered chalk to the solution of cream of tartar, until effervescence ceases. In the latter case, there will be an insoluble precipitate of tartrate of lime, from which the solution of the tartrate may be separated and evaporated, in order to obtain crystals. The following diagrams show the action in both cases.

180 or 1 equiv.	{	tart. acid 66	.....	114	tartrate of potash.
bi-tartrate		tart. acid 66	.....		
		potash 48	.....		
70 or 1 equiv.	{	potash 48	.....	114	tartrate of potash.
carb. potash		carb. acid 22	.....	22	carb. acid.
250		250		250	

1 equiv. of	{	tart. acid 66	.....	114	tartrate of potash.
bi-tartrate		tart. acid 66	.....		
		potash 48	.....		
1 equiv. of	{	lime 28	.....	94	tartrate of lime.
carb. lime		carb. acid 22	.....	22	carb. acid.
230		230		230	

*Magnesia, barytes, lime, acetate of lead, and nitrate of silver* decompose the tartrate of potash; so do all the *acids*, not excepting the carbonic, all *acidulous salts*, *tamarinds*, and *subacid fruits*, generally. These latter neutralise a part of the potash, sufficient to change the salt to a bi-tartrate. Hence it is manifestly improper, to administer soluble tartar in any acidulated drink, that could thus detach a part of its potash, because the mild cathartic effect would not be realised, in accordance with the expectation of the prescriber.

*Ferrocyanate of Potash.* This salt has been called the *triple prussiate*, and the *prussiate of potash*. It may be made by digesting the Prussian blue of the shops (ferrocyanate of the peroxyde of iron) in potash, until the alkali is neutralised, by which means the peroxyde of iron is set free, and a yellow liquid is formed, which yields crystals of the ferrocyanate of potash, by evaporation. This salt is made on a large scale, by igniting blood or other animal matters, as hoofs and horns, with potash and iron; by the mutual reaction of these matters, at high temperature,

the ferrocyanuret of potassium is generated, consisting of one equivalent of the radical of ferrocyanic acid, and two equivalents of potassium. That such is the product, is inferred from the circumstance, that by digesting the residue in water, we obtain ferrocyanate of potash.

This salt is soluble in three parts of water, at  $60^{\circ}$ , and in its own weight of boiling water, and the solution affords large, beautiful, yellow crystals. It is unaffected by the air, is blackened by a red heat, and finally decomposed. It is employed in the art of dyeing, and is also used in chemistry, as a reagent. Its solution is a very delicate and decisive test of the presence of iron, if it be, in part, in the state of peroxyde.

The composition of this complicated salt appears to be as follows, and 213 is consequently its equivalent number.

Cyanogen	78	or 3	equivalents.
Potassium	80	or 2	do.
Iron	28	or 1	do.
Hydrogen	3	or 3	do.
Oxygen	24	or 3	do.

---

213

---

*Chlorate of Potash.* This salt was formerly called *oxymuriate* and *hyperoxymuriate* of potash. It is colorless, soluble in 18 parts of cold, and in five parts of boiling water. It is susceptible of crystallisation; but the crystals are in form of scales, of a pearly lustre, and they contain no water of crystallisation. When exposed to a heat of  $400^{\circ}$ , they undergo the igneous fusion, and if the heat be now urged a little, effervescence ensues, and pure oxygen is evolved, in the way that was explained in the section on oxygen. It was there stated, that the chlorate of potash yielded purer oxygen than any known substance, because no foreign matter could operate to deteriorate it. This may be seen by the following diagram.

Chloric acid 76	{	oxygen	8	—	8	}	48 oxygen.		
		oxygen	8	—	8				
		oxygen	8	—	8				
		oxygen	8	—	8				
		oxygen	8	—	8				
Potash 48	{	chlorine	36			}	76 chloride of potassium.		
		oxygen	8	—	8				
		potassium	40						
<hr/>			124			<hr/>			
			124						
<hr/>			124			<hr/>			

Of course, 124 parts of the chlorate yield 48 of pure oxygen, and 76 parts of chloride of potassium remain in the retort.

This salt is made, by transmitting chlorine gas through a concentrated solution of pure potash, (or of the carbonate of potash,\*) until the alkali is quite neutralised. The solution, after boiling, contains nothing but muriate and chlorate of potash; it is gently evaporated till a pellicle forms on the surface, and then allowed to cool. The chlorate, for the most part, crystallises, while the muriate remains in solution. The crystals are to be washed in pure water, and then purified, by a second crystallisation.

The explanation of this play of affinities, is very simple. The water of the solution of potash is decomposed, or, at least, part of it; its oxygen goes to convert some of the chlorine into chloric acid, while its hydrogen joins other portions of the chlorine, to form muriatic acid, and these acids, united to the potash, give rise to the chlorate and muriate of potash.

The explosive nature of chlorine and oxygen, in the form of protoxyde and peroxyde of chlorine, has already been noticed. In the salt before us, chlorine is combined with five equivalents of oxygen, and its explosive power seems to depend on this large amount of the latter gas.

If one or two grains of the chlorate be rubbed smartly in a mortar, together with a little sulphur, (three or four grains,) loud explosions will be heard; the sulphur is said to combine with three equivalents of the oxygen of the chloric acid, and the chlorine escapes in form of gas.

Mix half a grain of phosphorus with a grain of chlorate of potash, in a small piece of paper, and fold it up. Then lay the mixture on a block of iron, and strike it with a smooth hammer, and the detonation will be violent. A larger quantity of materials is not safe, in the hands of a young operator, because the phosphorus is apt to be thrown out in sparks. In this case, the phosphorus combines with the oxygen of the chloric acid.

The *chloride of potassium*, remaining in the retort after oxygen is prepared from the chlorate of potash, is similar to that which is formed, when potassium is burned in chlorine gas. The same compound is formed, when muriatic acid is added to a solution of the carbonate of potash, and the solution thus obtained, evaporated and crystallised. The carbonic acid of the carbonate, is disengaged with effervescence, by the addition of the muriatic acid; and the hydrogen of this acid unites with the oxygen of the oxyde of potassium, forming water, while its

\* Mr Guthrie, who manufactures this article largely at Sacket's Harbor, says, the carbonate of potash answers sufficiently well.

chlorine joins the potassium, to give the chloride of potassium. If this chloride be dissolved in water, it is converted into muriate of potash. Part of the water of the solution is decomposed; its hydrogen unites with the chlorine of the chloride, and muriatic acid is the result, while its oxygen converts the potassium into potash, and this, in connexion with the muriatic acid, forms *muriate of potash*. This muriate was the *febrifuge salt of Sylvius*, but is, at present, scarcely known to the profession.

Some French chemists, especially Count Stadion and M. Serullas, have spoken of the *oxychlorate* or *perchlorate of potash*.

This salt may be formed, by mixing one part of powdered chlorate of potash, with three of strong sulphuric acid, and heating the mixture till it turns white, when the mass will consist of bi-sulphate and oxy or perchlorate of potash. It consists of 92 oxychloric acid and 48 potash. It is also formed, by the simple action of heat on the chlorate of potash, a part of which being decomposed, affords oxygen to that which is undecomposed, whence results the oxychlorate. Agreeably to Serullas, a heat little short of  $700^{\circ}$ , is requisite to effect this object.

*Hydriodate and Iodate of Potash.* These salts are sometimes formed by one and the same process, owing to the affinity which iodine exerts, in relation to the elements of water. Thus, when iodine is added to a concentrated, hot solution of pure potash, until the alkali is neutralised, the liquid will contain both the iodate and hydriodate of potash. The iodine decomposes the water, part of it uniting with the hydrogen and forming hydriodic acid, while the residue takes the oxygen, and forms iodic acid; and these, with the potash, give rise to the two distinct salts.

After the liquid above-mentioned has been evaporated to dryness, and quite cold, it is treated with alcohol, which takes up the hydriodate, and leaves the iodate behind.

The *iodate of potash* deflagrates on burning coals; heat decomposes it, evolving 23 parts of oxygen, while 77 of iodide of potassium remain. It requires  $13\frac{1}{2}$  parts of water, at  $60^{\circ}$ , for its solution. It is composed of iodic acid 164 parts, (one equivalent,) and potash 48, (one equivalent).

The iodate of potash may also be formed, by neutralising the iodic acid with potash; but the process is troublesome.

The *hydriodate of potash*, as formed in the process just detailed, may be separated from the alcohol, by a moderate heat. We may also convert the iodate, contained in the solution of potash, into hydriodate of potash, by passing through the liquid a stream of sulphuretted hydrogen gas. The hydrogen of this



gas decomposes the iodic acid of the iodate, combines with its oxygen, and then changes the iodine into hydriodic acid.

The best mode of obtaining this compound is, by the action of iron and iodine on water. A hydriodate of iron is thus produced, which must be decomposed by carbonate of potash; the products are carbonate of iron and hydriodate of potash. The whole is to be thrown on a filter, to separate the liquid hydriodate of potash, which must be subjected to evaporation.

To show the formation of the hydriodate of iron, 124 grains (one equivalent) of iodine, and 28 (one equivalent) of iron, may be mixed in a Florence flask, with two or three ounces of water, and heated with a spirit lamp, till the liquor becomes clear.

Every equivalent of iron and iodine decomposes one equivalent of water, and a reaction occurs, which is represented by the annexed diagram.

Before decomposition.		After decomposition.	
9 water	{hydrogen 1	125 hydriodic acid.	36 oxyde of iron.
	{oxygen 8		
Iodine	124		
Iron	28		
	<hr/> 161		<hr/> 161 hydriodate of iron.

If the solution of hydriodate of potash, obtained by the double decomposition of hydriodate of iron and carbonate of potash, be evaporated sufficiently, the hydrogen of the hydriodate will unite with the oxygen of the potash, forming water, while the iodine and potassium will unite to make *iodide of potassium*. We see, therefore, the reason why this iodide is so generally confounded with the hydriodate. The same principle of chemical action applies here, that was noticed in relation to the chloride of potassium and the muriate of potash.

The decomposition of the iodate of potash, by heat, is another source of the iodide of potassium. The iodic acid and potash both lose their oxygen, while the iodine and potassium, being thus set free, join to form the iodide.

The hydriodate of potash is the most important of all the hydriodates that are known. It is soluble in two-thirds its weight of water at 60°, and in much less at 212°; it is also very soluble in alcohol. It deliquesces in a humid atmosphere, and should, therefore, be kept in tight vessels. It is found native, especially in mineral waters, where it is detected by the agents named, when treating of hydriodic acid. The solutions of this salt dissolve, readily, a large quantity of iodine,

and they thus acquire a brown color. The solution, thus made, is called the *ioduretted hydriodate of potash*; and it is now the most popular mode of using iodine. It may be made, by adding ten grains of iodine to thirty-six grains of hydriodate of potash, in an ounce of distilled water; from six to ten drops are given in syrup and water, three or four times a day. The same solution, but of greater strength, is used as a lotion, in scrophulous ophthalmia.

An overdose of the hydriodate, or the ioduretted hydriodate, occasions similar symptoms to those induced by iodine, and the treatment is the same, in all the cases.

Several other salts of potash might be introduced here, if they were regarded as being sufficiently important. There are others, whose formation and properties cannot, with propriety, be noticed, until we examine the acids which enter into their composition; such are the arsenite and chromate of potash, whose acids are of metallic origin.

---

## CHAPTER II.

### Of Sodium.

THIS metal, which is the base of soda, was first obtained by Sir H. Davy, in 1807, shortly after he procured potassium. The agent he employed to effect this object, was galvanism; and in the decomposition thus induced, the oxygen of the soda always went to the positive, and the sodium to the negative, pole of the battery. This metal may also be procured in the other ways mentioned for preparing potassium, and also by mixing chloride of sodium with potassium, and distilling from an iron retort, when the potassium will combine with the chlorine, and the sodium be volatilised.

In all its chemical relations, sodium has a close resemblance to potassium, and we may exhibit with the former, the various experiments detailed in connexion with the latter. It does not burn, however, as potassium does, when thrown on cold water; but if the water be made hot, combustion readily takes place.

In point of color and metallic lustre, sodium (as well as potassium) resembles silver. At ordinary temperatures, it is so soft, that it may be formed into leaves by the pressure of the fingers. At  $200^{\circ}$  it fuses, and at a full red heat rises in vapor. Its specific gravity is a little greater than that of potassium, but still below that of water; it is 0.972. The equivalent, or com-

binning number of sodium is 24. It soon tarnishes on exposure to the air, though less rapidly than potassium.

The burning of sodium on hot water, or its action on cold water, gives proof that it is really the base of an alkali, as may be seen by the addition of a little turmeric infusion, or the use of the turmeric paper.

Sodium combines readily with potassium, and the compounds vary in their properties, according to the proportions of the constituents. By a small quantity of sodium, potassium is rendered fluid at ordinary temperatures, and its specific gravity is lessened. Eight parts of potassium and one of sodium, form a compound that swims on naphtha, and is fluid at 60°, F. Three parts of sodium and one of potassium, make a compound that is fluid, at common temperatures. A little potassium destroys the ductility of sodium, and renders it very brittle and soft. Since the equivalent of potassium is to that of sodium, as five to three, it will require the former quantity of potassium, to eliminate or detach the latter quantity of sodium from its chloride. The attractions of potassium for all substances that have been examined, are stronger than those of sodium.

There are two compounds of sodium and oxygen, called the protoxyde and peroxyde. The first is the article usually called *soda*; it is a grayish-white solid, difficult of fusion, and may be made by burning sodium in dry atmospheric air. It is also formed, when sodium is oxydised by water, and its composition may be known, by collecting the hydrogen which is then disengaged. Every equivalent of water, viz. 9, is decomposed by 24 or one equivalent of sodium; the oxygen of the water, 8, combines with the metal, while one equivalent of hydrogen is evolved. The protoxyde is, consequently, composed of 24 sodium and 8 oxygen.

The impure, commercial substance, named *barilla*, is obtained from the ashes of the plant called *salsola soda*. *Kelp*, a much coarser article, is procured from the ashes of ordinary seaweeds; it contains from 2 to 5 per cent. of real soda, while *barilla* yields about 20. It is affirmed, that the soda plants which grow along the sea-coast will not yield a particle of soda, if transplanted to an inland situation. This leads naturally to the inference, that these plants have the power of decomposing the muriate of soda, existing in sea-water, and appropriating the soda to their own use.

Soda is the substance formerly known, by the term *mineral alkali*, because under the name of *natron*, it was found in seams or crusts in minerals. The article thus designated, was not pure soda, but a compound of soda and carbonic acid.

The *peroxyde* of sodium is formed, by heating the metal in an excess of pure oxygen. It is an orange-colored substance, which is changed by the action of water into oxygen and soda. According to Gay Lussac and Thenard, it is a compound of two equivalents of sodium and three of oxygen, or of sodium one equivalent, and oxygen one and a half.

Soda is usually obtained from the carbonate of soda, by following the process described for the preparation of potash, from its carbonate. With water soda forms a solid hydrate, easily fusible by heat; it is highly caustic, soluble in water and alcohol, possessing decided alkaline properties, and is analogous, in most respects, to potash. The solid hydrate is composed of 32 parts or one equivalent of soda, and 9 parts or one equivalent of water. Ure mentions another compound of oxygen and sodium, which he calls the *suboxyde* of sodium. When the metal is kept for some time in a small quantity of moist air, or when it is heated, in excess, with hydrate of soda, a dark grayish substance is formed, that is more inflammable than sodium, and which yields hydrogen by its action upon water. The best modes of distinguishing soda from other alkaline bases, are the following. It yields, by union with sulphuric acid, a salt which may easily be recognised as the sulphate of soda or Glauber's salt. All its salts are soluble in water, and are not precipitated by any known reagent. If we expose its salts to the flame of a blowpipe, suspended by a platina wire, a rich, yellow color will be perceptible. Soda is the basis of common salt, hard soaps, and plate and crown glass.

*Chloride of Sodium.* This is composed of one equivalent of chlorine, 36, and one of sodium, 24. It may be made by burning sodium in chlorine, or by heating it in muriatic acid gas. In the latter case, it effects a decomposition, taking chlorine from the muriatic acid and setting its hydrogen at liberty. If a solution of common salt (muriate of soda) be evaporated, crystals of the chloride are deposited; for this salt, like the muriate of potash, can exist, as such, only in solution, and is converted into a chloride by expelling the water, so as to reduce the whole to crystals. We are, therefore, to regard the solution of common salt as a true muriate of soda, and the same solution reduced by evaporation to the crystallised state, as a chloride of sodium. When water is added to the chloride, that liquid is decomposed; its oxygen joins the sodium to form the protoxyde of soda, while its hydrogen unites to the chlorine to form muriatic acid, and the muriatic acid combines with the protoxyde of soda, to make the muriate of soda. Sodium has a much stronger attraction for chlorine than for oxygen, and

hence, soda or its hydrate is decomposed by chlorine, and a chloride of sodium is formed.

Although the names chloride of sodium, and muriate of soda, are, chemically, distinct, yet they are generally used to denote the well known article, common salt. This substance exists in the mineral kingdom, in immense beds, to the extent of several hundred miles in length, and when obtained thence, is called *rock salt*. What is termed *fish salt*, *stoved salt*, *sea salt*, and *bay salt*, is procured from sea-water, salt-licks, and springs, by evaporation. This process is carried on, spontaneously, in warm climates; the water being placed in large reservoirs, for that purpose. In more northern regions, the water is dissipated by boiling in large pans, and sometimes advantage is previously taken of the very cold weather, to freeze part of the water, which thus leaves a much stronger brine. Much larger and purer crystals are obtained, when the sea-water is evaporated spontaneously, than when considerable heat is employed. In the latter case, other salts, as the muriate of magnesia,\* combine with the muriate of soda, rendering it more deliquescent, and thus impairing its antiseptic properties.

We may easily determine, whether magnesia be really present in a solution of common salt, by adding a solution of carbonate of soda; if magnesia be present, a white precipitate of carbonate of magnesia will be seen, otherwise the solution will remain as colorless as before.

The crystals of dry muriate of soda, or chloride of sodium, contain no water of crystallisation, but decrepitate when heated.

The pure chloride has an agreeably saline taste, fuses at a red heat, and becomes a transparent, brittle mass, on cooling. It deliquesces slightly, in a moist atmosphere, but undergoes no change when the air is quite dry. In pure alcohol, it is wholly insoluble. It requires twice its weight of water, at  $60^{\circ}$ , for solution, and it does not dissolve more freely when the water is heated. Sulphuric acid decomposes the muriate of soda, with evolution of muriatic acid gas, and sulphate of soda is the product. If anhydrous sulphuric acid were added to the chloride of sodium, there could not be a disengagement of muriatic acid gas; but if water be present in the acid, its decomposition furnishes hydrogen, which joins the chlorine of the chloride to

\* This muriate may be separated from common salt by exposure to a gentle heat, in a reverberatory furnace. It parts with its acid, at a heat a little over that of boiling water, after which, if the salt be dissolved in water, the magnesia will be precipitated.



form muriatic acid, and this escapes in the state of gas; the oxygen of the water joins the sodium, making an oxyde of sodium, or soda, which combines with the sulphuric acid, giving rise to sulphate of soda.

The uses of chloride of sodium, or muriate of soda, are various. Its employment in seasoning food and preserving meat from putrefaction, is well known, though on what precise principle it acts, in the latter case, is not well understood. It is used for various purposes in the arts, especially for making muriatic acid and chlorine gas.

Common salt has long been used in the practice of medicine, especially in the treatment of worms. It was a favorite remedy with Dr Rush, who administered it, colored with cochineal to prevent the common people from knowing it; not because he wished to keep them in ignorance, but on account of their prejudice against the use of simple remedies.\*

Chloride of sodium is sometimes confounded with *chloride of soda*, but they are different articles. The latter seems to offer a departure from the analogy, generally observed, between the compounds of oxygen, and those of chlorine, iodine, &c. The chloride of soda is, evidently, a chloride of an oxyde, and it is certain, we have no instance of an oxyde of a chloride. For this reason, it was at first doubted, that chlorine could combine, chemically, with either lime or soda, and I do not believe that a true chemical union exists. It is most likely, that the lime and soda retain the chlorine, pretty much as charcoal retains large bulks of various gases, or as water holds in solution ammoniacal gas. Dr Ure speaks of the *absorption* of chlorine by the hydrate of lime, and it certainly remains to be shown, that any other state of union exists in these chlorides, than that of simple mixture. The evidence in favor of the chemical composition of atmospheric air is, in my view, far stronger than any that has yet been adduced, to prove that chloride of soda is a true chemical compound.

Pure chloride of soda is prepared, by passing a current of chlorine gas into a cold and dilute solution of caustic soda, until no more gas can be taken up. Common carbonate of soda may be used in place of the pure alkali; but considerable excess of chlorine must be employed, in order to displace the whole of the carbonic acid.

\* The ancient laws of Holland ordained, that criminals should be kept on bread alone, *unmixed with salt*, as the *severest* punishment they could inflict upon them, in their moist climate. The effect was horrible; the wretched beings were *devoured by worms*, engendered in their own stomachs, from want of salt. See Paris's Pharmacologia, vol. ii.

*Labarraque's disinfecting soda-liquid* is a chloride of soda, and has been examined by Mr Phillips and Mr Faraday, who give the following formula for its preparation. Two thousand eight hundred grains of crystallised carbonate of soda are dissolved in little more than a pint of water, and chlorine gas\* is passed through the solution in Woulfe's apparatus, taking care, first, to transmit it through a bottle of pure water, to remove any muriatic acid that may be mixed with it. No carbonic acid is disengaged, and though the solution contains chloride of soda, its precise composition has not been accurately determined. If kept in open vessels, it is in a great measure deprived of its free chlorine, and chlorate and muriate of soda are formed. The water of the liquid, being in part decomposed, its hydrogen changes part of the chlorine to muriatic acid, while its oxygen converts another portion into chloric acid, and these, combining with the soda, give rise to the chlorate and muriate.

M. Fayen proposes to make the chloride of soda, by mixing chloride of lime with carbonate of soda. He says the chloride of lime is decomposed by the soda, and that the chloride of soda, thus made, is of uniform strength. But in whatever way it is prepared, the properties of the article are the same. The chlorine is held in connexion with the soda by so feeble a tie, that it is easily set at liberty. As proof of this, it always emits the natural odor of chlorine, no matter how carefully it may have been prepared, and it possesses the peculiar bleaching powers of chlorine, in a high degree.

It is stated, in some of the books, that when the chloride of soda undergoes decomposition, by exposure to the air in open vessels, it is the carbonic acid of the atmosphere that accomplishes this result. But it is not very easy to reconcile this statement with the assertion of Mr Faraday, that in partially charging a solution of carbonate of soda with chlorine, no carbonic acid is evolved, and yet chloride of soda is formed. Now it is plain, that if there be any chloride produced, there must be some decomposition of the carbonate, and although its carbonic acid may not show itself in effervescence, it must be present, in some state. And if its proximity to the chloride of soda, in its atmospherical association, be capable of decomposing the chloride, it should be equally operative, when actually suspended in the same water, which holds the chloride in solution. Is there not something to be revealed, in order to the full elucidation of

\* They direct the chlorine, for this use, to be prepared from 967 grains of common salt, 750 of peroxyde of manganese, and 960 of sulphuric acid, diluted with an equal quantity of water.

this subject? To us it appears, that much yet remains to be determined by careful experiments.

Chloride of soda is useful as a bleaching agent, and for all other purposes to which chlorine has been applied. It removes the offensive odor from sewers, privies, animal substances in a state of putrefaction, and foul ulcers, that may not be actually gangrenous. It is also used to disinfect clothing that may have been exposed to pestilential effluvia, for purifying sick chambers, &c. The extrication of the chlorine is so gradual that its irrespirable quality is not perceived. The chlorine of the chloride appears to neutralise noxious exhalations, by seizing upon the hydrogen, which always enters into their constitution, and thus destroying the exhalations themselves.

As the chloride of soda is a more expensive article than the chloride of lime, it is not very much employed in this country. *Labarraque's liquid* is certainly a very convenient article, being just in the state for use at any moment, when it may be required.

The compounds resulting from the union of sodium with iodine, sulphur, and phosphorus, being prepared in the same manner as the compounds of potassium and those simple, non-metallic bodies, and their properties being nearly the same, it is not necessary to give a particular account of them. Sodium does not unite with hydrogen, as potassium does.

The salts of soda have the same general properties as those of potash, but are, for the most part, more soluble in water, and all of them may be decomposed by this alkali. To show the greater affinity of potash for the acids, pour a strong solution of this alkali into a saturated solution of the sulphate of soda; crystals of sulphate of potash are soon formed, and soda remains in solution.

The *sulphate of soda*, or Glauber's salt, is prepared from the residuum of the preparation of muriatic acid, in the same manner as the sulphate of potash is procured from the residuum of the distillation of nitric acid. It crystallises in prisms which are soluble in three parts of water, fuse on exposure to heat, and part with their water of crystallisation. It is occasionally found at the earth's surface, and is frequently met with in mineral springs. It may be made by the direct union of its component parts, viz. sulphuric acid and soda; but it is manufactured very largely in the process for obtaining sal ammoniac, or muriate of ammonia. The sulphate of ammonia, an article existing in coal soot, is decomposed by common salt or muriate of soda, and the two resulting compounds are muriate of ammonia and sulphate of soda.

This salt, so well known in domestic practice, has a cooling,

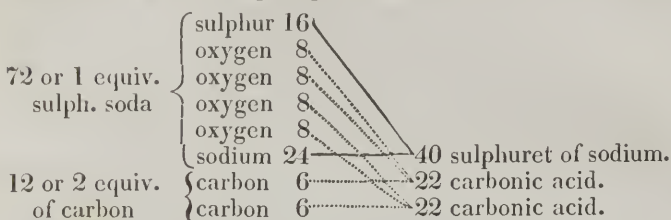
saline, and bitter taste; and these properties seem to distinguish it sufficiently from saltpetre, with which it has been sometimes confounded. We have noticed this subject, under the article, nitrate of potash, to which the reader is referred. The crystals of sulphate of soda effloresce rapidly, or are reduced to powder, by exposure to the air; and, according to Berzelius, they are composed of seventy-two parts or one equivalent of neutral sulphate, and ninety, or ten equivalents of water. Water at  $60^{\circ}$ , Fahrenheit, dissolves one third of its weight of the crystallised salt; at  $72^{\circ}$ , nearly half its weight; twice its weight at  $90^{\circ}$ ; and three times its weight at  $106^{\circ}$ . If the heat be elevated beyond this point, a portion of the salt will be deposited, because less soluble than at  $106^{\circ}$ . If a saturated solution at  $106^{\circ}$ , is evaporated at a higher temperature, the salt separates in opaque, prismatic crystals, which are anhydrous. It is quite insoluble in alcohol.

The fine powder of sulphate of soda, resulting from exposure of the crystals to the air, is supposed, by the ignorant, to be of no value. In this effloresced state, it is equal, in point of medicinal activity, to double the weight of the crystallised salt. The exact quantity of water, required to dissolve the effloresced salt, has not been given in any of the books, and I have not had leisure to make the experiment. Some have supposed it would require a quantity equal to its original water of crystallisation, but there seems to be no good reason for this conjecture. By dissolving it, we do not restore it to its original form, and we know further, that the crystals of this salt do not always contain the same relative quantity of water. Usually they are composed of 40 dry acid, 32 soda, and 90 of water; but, a saturated solution of sulphate of soda, corked up in an air-tight vessel, gradually deposits crystals, at first quite transparent, but which become opaque, when the bottle is opened, and these, according to Mr Faraday, contain forty parts dry acid, thirty-two soda, and seventy-two of water, or eighteen less of water than the common crystals. Moreover, the effloresced salt still retains a portion of the water of crystallisation, and is susceptible of the watery fusion. The *incompatibles* of sulphate of soda are, the *muriates of ammonia*, *barytes*, and *lime*, *nitrate of silver*, *acetate and subacetate of lead*; mixed with either of which, it occasions double decomposition.

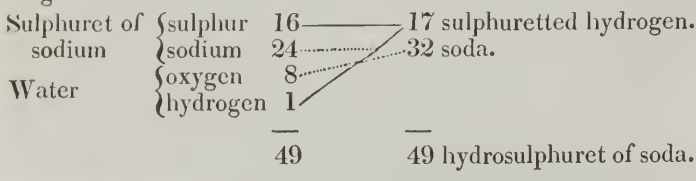
The nauseous taste of Glauber's salt may be much obviated, by adding a small quantity of lemon juice, or cream of tartar, neither of which occasions decomposition.

Sulphate of soda is called, in the old books, by various names, as *vitriolated natron*, *Glauber's purging salt*, *vitriolated soda*, *sal mirabile*, &c.

The sulphate of soda, exposed to heat in a common crucible, mixed with charcoal, is decomposed, and sulphuret of sodium is formed, as the following diagram will show.—



When the sulphuret of sodium, thus formed, is acted on by water, a hydrosulphuret of soda is the result, as the following diagram evinces.



*Bi-sulphate or supersulphate of soda* may be formed, by dissolving the sulphate in sulphuric acid, and evaporating the solution.

*Sulphite of soda* is a compound of sulphurous acid and soda.

There are three compounds of soda with carbonic acid, mentioned in the books, viz. the *carbonate*, *bi-carbonate*, and *sesqui-carbonate*.

The *carbonate* of soda has been called a *subcarbonate*, because the alkali was supposed to be in excess; whereas, there is precisely one equivalent of acid, united to one of alkali, in this salt. It is composed of carbonic acid 22, soda 32, and 90 of water, when in the form of crystals, though the proportion of water may vary, according to the mode of obtaining the carbonate.

The *carbonate of soda* of commerce is obtained by digesting kelp or barilla in water, and crystallising the solution; or by soaking saw-dust or other carbonaceous matter, in a strong solution of the hydrosulphuret of soda and exposing the mixture to heat, in a reverberatory furnace. The water is soon dissipated, the sulphuretted hydrogen decomposed and expelled, and the soda combines with carbonic acid, formed by the combustion of the fuel or of the materials with which the solution is mixed.

Carbonate of soda is easily obtained in large crystals, from its solution in water. They are soluble in two parts of cold water, and in less than their own weight of boiling water; on exposure to the air they effloresce, and fuse when heated, losing their



water of crystallisation, the quantity of which has been found to vary, according to the temperature at which the crystals are formed. Its solution turns the vegetable colors to a green.

The taste of this salt, though mild, is decidedly alkaliescent, and being much less nauseous than the carbonate of potash, is preferable for medicinal purposes. Fourcroy maintains the preference of soda on another ground, viz. that it has more analogy with the animal economy, being always found in the human fluids, while potash is rarely if ever met with.

The incompatibles of the carbonate of soda are the same with those mentioned under the article, carbonate of potash.

The carbonate of soda acts, chemically, in reducing the quantity or preventing the formation of the red sediment in urine, as well as in removing acidity of the stomach.

The *dry carbonate of soda* is made, by exposing a pound of the carbonate, in a clean iron vessel, to a boiling heat, and stirring constantly with an iron rod, until it is perfectly dry, after which it must be reduced to powder. The heat employed drives off about 25 per cent. of the water of crystallisation, and fits the article for use in form of pills. If made into pills in the undried state, they would fall to pieces, owing to the salt rapidly efflorescing, or giving out its water to the air.

*Bi-carbonate of soda* is generally called the carbonate, but its real composition shows the correctness of the prefix *bi*. Exclusive of its water of crystallisation, it contains 32 or one equivalent of soda, and 44 or two equivalents of carbonic acid.

Any process, by which we can double the quantity of carbonic acid, existing in the carbonate, will give us the bi-salt. It may be prepared, however, in the manner detailed for making the bi-carbonate of potash.

The bi-carbonate of soda is evidently alkaline, but much more pleasant than the carbonate; it is far less soluble, requiring about ten times its weight of water, at 60°, for solution. It is decomposed partially at 212°, and is changed into the carbonate, by a red heat. The real *soda water* is merely water holding in solution carbonate of soda, and containing an excess of carbonic acid gas. But in very many instances, what is sold for soda water, does not contain a particle of soda, and is nothing more than water highly impregnated with carbonic acid gas. The latter, though a very pleasant beverage, has comparatively little medical efficacy.

The *sodaic* powders, now in such general use, as a substitute for soda water, consist of bi-carbonate of soda and tartaric acid. About thirty grains of each of these articles are put up in separate papers, which are of different colors, to designate their

contents. The acid and alkali are dissolved in two separate tumblers, containing about a half pint of water; and on mixing the contents of the tumblers, an agreeable effervescing draught is made, which should be swallowed immediately. The tartaric acid combines with the soda of the bi-carbonate, forming tartrate of soda, while the carbonic acid escapes.

The *sesqui-carbonate* of soda is said to contain one equivalent of soda, and one and a half of acid, with two equivalents of water. It is said to have been found on the banks of the lakes of soda, in the province of Sukena, in Africa, whence it has been exported, under the name of *trona*.

Its existence, under ordinary circumstances, is purely accidental, and depends on the want of regularity in applying heat to the bi-carbonate. If the heat be above  $120^{\circ}$ , part of the bi-carbonate will lose one equivalent of acid and another part will lose a half equivalent, thus giving rise to the sesqui-carbonate.

*Phosphate of soda* is prepared, by adding carbonate of soda to a solution of the bi-phosphate of lime. The excess of phosphoric acid in the bi-phosphate, combines with the soda of the carbonate, and remains in solution, disengaging carbonic acid with effervescence, while the phosphate of lime, having now lost its excess of acid, becomes insoluble and is precipitated. In the following diagram, representing the action that takes place, two equivalents of acid are supposed to be combined with one of lime in the bi-phosphate.

	Before decomposition.		After decomposition.	
Carbonate of soda	{ carb. ac.	22	22	carbonic acid.
	{ soda	32	60	phosphate of soda.
Bi-phosphate of lime	{ phosph. ac.	28		
	{ phosph. ac.	28		
	{ lime	28	56	phosphate of lime.

The bi-phosphate of lime, however, is generally prepared with a much greater excess of acid, and every additional equivalent which it contains, enables it to decompose another equivalent of the carbonate of soda. The latter ought to be added, till the solution can render the test-paper green, as a slight excess of alkali favors the crystallisation. The solution is filtered to separate the phosphate of lime, and evaporated afterwards till a pellicle appears on its surface, when it may be set aside to crystallise.

Phosphate of soda is soluble in four parts of cold, and two of boiling water. It has a saline taste, unaccompanied by the disagreeable bitterness of sulphates of magnesia and soda. By exposure to a moderate heat, it loses most of its water of crystallisation, and its properties are considerably altered.

It is the *sal perlatum* of some old writers, and was introduced into pharmacy by Dr Pearson. It is also called *soda phosphorata*, and as it occurs in the shops, is always adulterated with sulphate of soda. Phosphate of soda has been used as a cathartic, and was known in that character by the name of the *tasteless purging salt*. It is used also, as a test, to discover the presence of magnesia, in mineral waters. A *bi-phosphate of soda* is named in some of the books, but it is of too little importance to claim our particular notice.

*Tartrate of Potash and Soda.* This is a triple salt, containing two bases and one acid. It is called by Paris, Brande, and others, *soda tartarisata*, a name which is not sufficiently full, to indicate its composition. *Salt of Seignette* and *Rochelle salt* are the names first given to this article; and the latter is still in use, among the common people. It is soluble in five parts of water at  $50^{\circ}$ , and is very slightly efflorescent.

The *tartrate of potash and soda* is prepared, by adding the bi-tartrate of potash, (cream of tartar,) in fine powder, to a solution of the carbonate of soda in boiling water, as long as any effervescence takes place. Three hundred grains of the carbonate, dissolved in ten or twelve parts of water, will be sufficient to show the process; and for this, about 400 grains of cream of tartar are required. The excess of acid, which it contains, unites with the soda of the carbonate, and the carbonic acid is disengaged with effervescence; the tartrate of soda remains in combination with the tartrate of potash, and the solution being evaporated to a pellicle and set aside to cool, yields regular crystals. The following diagram gives a view of the reaction.

Before decomposition.				After decomposition.	
Carbonate of	{	carbonic acid	22	-----	22 carbonic acid.
soda		soda	32	-----	98 } tart. of potash and soda.
Bi-tartrate of	{	tartaric acid	66	-----	
		tartaric acid	66	-----	
		potash	48	-----	
				-----	114

The diagram represents this salt as a compound of 98 parts of tartrate of soda and 114 of tartrate of potash; but there are also in its composition, eight equivalents of water, making its actual equivalent 284.

It is a very pleasant, gentle cathartic, and is a favorite medicine with the French. It enters into the composition of the effervescent aperient commonly sold under the title of *Seidlitz powders*. These consist of two different powders; the one contained in a white paper consists of one hundred and twenty grains of the tartrate of potash and soda, with forty grains of bi-

carbonate of soda; the other, in a blue paper, consists of thirty-five grains of tartaric acid. The contents of the white paper are to be dissolved in half a pint of spring water, to which the tartaric acid is to be added; effervescence immediately occurs, and the draught is to be instantly swallowed. The excess of acid renders the mixture more agreeable, and does not diminish its efficacy as a purgative.

*Nitrate of soda* or *cubic nitre* is obtained, by saturating the mother-waters of the saltpetre works, with carbonate of soda instead of wood ashes; or, by saturating carbonate of soda with nitric acid, evaporating and crystallising. Prout recommends this salt to fire-workers, as an economical substitute for the nitrate of potash, as it burns about three times as long.

*Succinate of soda* is employed only as a reagent, to detect the presence of iron in mineral waters. The succinate is decomposed by the iron and a succinate of iron is formed.

This salt is prepared, by adding succinic acid to carbonate of soda water, until the liquor contains a slight excess of acid, beyond what is necessary for the saturation of the alkali.

The *bi-borate of soda*, or *borax*, is a native production, existing in the water of some lakes in Thibet and Persia. The crude borax met with in commerce, and known by the name of *tincal*, is refined by solution and crystallisation. It is often termed subborate of soda, as its solution turns the vegetable blues to a green; two equivalents of boracic acid not being sufficient to neutralise one equivalent of soda, so as to destroy its alkaline properties.

This salt crystallises in prisms, which are soluble in six parts of boiling and twenty of cold water. When exposed to the air, it effloresces; and heat first drives off the water of crystallisation, and then fuses the mass, forming a vitreous, transparent substance, called *glass of borax*. According to the analysis made by Dr Thomson, the crystals of borax contain 48 parts or two equivalents of boracic acid, 32, or one equivalent of soda, and 72, or eight equivalents of water.

The chief use of borax is as a flux, and to prepare boracic acid. In the practice of medicine, it is occasionally employed in form of powder, mixed with honey, or dissolved in astringent decoctions. Its incompatibles are the *acids*, especially the mineral; also, *potash*, the *sulphates* and *muriales* of the *earths*, and of *ammonia*.

The *calcined borax* of the books is the bi-borate of soda, heated so as to lose its water of crystallisation, without carrying the heat any further. It is then a porous and friable mass, somewhat resembling burnt alum.

## CHAPTER III.

## Of Ammonium.

ON the authority of Berzelius, I add ammonium to the list of alkalifiable metals. For although there are, evidently, some facts connected with this substance that are not easily reconciled with our views of metals, in general, yet the experiments of the Swedish philosopher forbid us to deny it a place in the list of metallic bodies. He affirms that it is not a simple, but a compound substance.\*

By several methods, detailed in the late work of Berzelius, an amalgam was formed between the metallic base of ammonia and mercury, of such an unequivocal character, as to satisfy all who witnessed the experiments. The amalgam is spoken of, as being quite as soft as butter, and occupying five or six times as large a volume as the original materials. At length it is changed into ammonia, with evolution of hydrogen, and the mercury regains its original bulk.

In this view of the metallic base of ammonia, and recalling its acknowledged gaseous composition to our recollection, we are forced to the conclusion, that either the nitrogen or hydrogen which enters its formation, or both, must be of metallic origin. But as all suggestions of this kind are at best hypothetical, and, of course, not suited to an elementary work, we shall pass to the examination of the article usually called ammonia, and the various compounds of which it makes a part.

*Of Ammonia.* Whatever may be the exact nature of the metallic base of ammonia, it is agreed by all chemists, that this alkaline matter is composed of nitrogen and hydrogen; hence it is affirmed, that whenever an attempt is made to decompose ammonia, without the aid of mercury, it is resolved into the above-named gases, in the proportion of 14 nitrogen to 3 hydrogen, making the equivalent of ammonia 17.

Pure ammonia is a colorless gas, which has a strong, pungent odor, and acts powerfully on the eyes and nose. It is quite irrespirable, and quickly extinguishes burning tapers, but if considerably diluted with air, it may be taken into the lungs with safety. Although it extinguishes lighted tapers, the flame is, at first, a little enlarged, and a small jet of the gas actually burns in an atmosphere of oxygen.

Dr Priestley was the first to point out the existence of ammo-

\* This opinion is not sufficiently established, by conclusive experiment.



nia, but he called it *alkaline air*. It is sometimes called the *volatile alkali*, because of its tendency to escape. It is also called *spirits of hartshorn*, because it was formerly made from the horns of the hart and other animals, and is an animal product to a great extent.

A very convenient mode of preparing ammonia or ammoniacal gas, is to expose the common spirits of hartshorn or liquid ammonia, to heat. The ammonia quickly and rapidly passes over in form of gas, and may be collected in suitable glass vessels placed on a mercurial trough.

But this mode is rather in anticipation, and we therefore say, that the gaseous ammonia may be procured, by decomposing the sal ammoniac of the shops, by means of quick-lime. If equal parts of these substances be reduced to fine powder, placed in a glass retort, and exposed to the action of a spirit lamp, the gas may be readily collected over mercury. The nature of the action will be seen by the following diagram.

54 or 1 equiv.	{ ammonia	17	—————	17 ammonia.
of muriate	{ hydrogen	1	.....	9 water.
of ammonia	{ chlorine	36	.....	
28 or 1 equiv.	{ oxygen	8	.....	
of lime	{ calcium	20	—————	56 chloride of calcium.
—		—		—
82		82		82
—		—		—

In this diagram, it is seen, that sal ammoniac is a compound of muriatic acid and ammonia, and that the acid is composed of chlorine and hydrogen; it is also perceived that the lime is a compound of oxygen and calcium. The chlorine of the acid combines with the calcium of the lime, to form chloride of calcium, while the hydrogen of the acid joins the oxygen of the lime to form water, leaving the ammonia entirely detached, and ready to take on its native, gaseous form.

When large quantities of the above materials are employed, and the gas is passed into water in a close vessel, the *liquid ammonia*, or *water of ammonia*, of the shops, is formed. It is also made, by decomposing animal matters by heat in close vessels, and condensing the product in water.

The specific gravity of ammoniacal gas is 0.5902, while that of the liquid ammonia varies according to the quantity of water employed to condense the gas. Water at 50°, (the barometer standing at 29.8,) absorbs 670 times its volume of the gas, and the solution will have a specific gravity equal to 0.875.

Ammonia possesses all the qualities of an alkali, affecting the vegetable colors in the same manner as potash and soda, and forming salts which are decomposed or volatilised by heat. In noticing its effects on the vegetable colors, it should be remembered, that if turmeric paper be immersed in it, and then laid aside for some hours, the yellow color will appear unchanged, owing to the volatile nature of the alkali.

Alcohol has a strong affinity for this gas, and condenses a large quantity of it. Hence the various alcoholic preparations of ammonia, noticed in the books, of which the *alcohol ammoniatum* is the basis. The *spiritus ammoniæ aromaticus* or the *volatile aromatic spirit*, is merely the ammoniated alcohol, holding several essential oils in solution. The *fetid spirit of ammonia* is a solution of the fetid matter of assafœtida in the ammoniated alcohol.

To see the great affinity that water has for ammonia, take a tube or bottle filled with this gas over mercury, and open it under water in a basin; the ammonia will be instantly condensed by the water, which will rise into the tube and fill it completely, if the gas has been quite pure. Dr Thomson states, that water can take up 780 times its volume of ammonia at common temperatures.

When this gas is mixed with oxygen and inflamed, the mixture must be made over mercury in a strong jar or detonating bottle; every two measures of ammoniacal gas (one equivalent) require a measure and a half of oxygen (three equivalents) for combustion, the hydrogen and the oxygen combining together and forming water, while the nitrogen is disengaged.

Ammonia is decomposed by chlorine, which, combining with its hydrogen, forms muriatic acid, and disengages nitrogen; if an excess of ammonia is used, the muriatic acid combines with it, and forms muriate of ammonia. The mixture may be made over the mercurial trough, or a bottle full of chlorine may be connected with a bottle full of ammoniacal gas by a glass tube passing through a cork that fits them both, placing the bottle containing the chlorine uppermost. A lambent flame accompanies the action; three measures of chlorine are required for the complete decomposition of two measures of ammonia.

Ammonia is easily recognised by its odor and by the white fumes of muriate of ammonia which are formed, when a glass rod, dipped in muriatic acid, is brought in contact with it. When its solution in water is too dilute to give any perceptible smell, it may in general be detected by the deep blue color which it produces with solutions of salts of copper.

Ammonia, whether in form of gas or in the shape of liquid ammonia, is capable of acting as a poison. When the gas has been accidentally inhaled in considerable quantities, the patient should be quickly removed to a free and pure air, be well washed with vinegar and water, and be made to swallow a portion of the same, mixed fluid.

When the liquid ammonia has been swallowed, a course somewhat similar will be proper. On the principle of neutralising the alkali, we should think a weak acid would be highly proper, and none more so, than the acetic, moderately diluted. But recently, M. Chereau, a French physician, has stated that a fixed oil is the best antidote for all the alkalis, as well as for their carbonates. He says, they act partly by rendering the vomiting free and easy, and partly by converting the alkali into a soap.

If ammoniacal gas be passed through charcoal, ignited in an earthen tube, prussic or hydrocyanic acid will be formed. In this process, the ammoniacal gas is decomposed; its nitrogen, combining with the carbon, forms cyanogen, and this, joined to a portion of the hydrogen of the ammonia, gives rise to hydrocyanic acid.

Iodine exerts a decomposing energy on ammonia. It unites with the nitrogen of the ammonia, and forms the explosive iodide of nitrogen, formerly noticed. With the hydrogen of the ammonia, it forms hydriodic acid, and if an excess of alkali be present, it will combine with this acid, producing the *hydriodate of ammonia*.

Sulphuretted hydrogen readily combines with ammonia, and the resulting compound is called *hydrosulphuret of ammonia*. This may be prepared, by passing sulphuretted hydrogen and ammoniacal gas into a glass balloon, or common Woulfe's bottle, through two distinct apertures, and at a low temperature. As soon as the gases meet, transparent white, or whitish-yellow crystals are formed. If a mere solution of the hydrosulphuret is wanted, either for medicine or analysis, it may be prepared by passing a stream of sulphuretted hydrogen gas through liquid ammonia, until it is completely saturated.

The hydrosulphuret of ammonia has been employed as a lotion, in the treatment of cutaneous affections; it is also used as a test, in all cases to which sulphuretted hydrogen is adapted.

Potassium and sodium exert a very curious agency on gaseous ammonia. When fused in the gas, an olive-green substance, of easy fusion, is formed, consisting of the metal, nitrogen, and ammonia. Supposing that potassium is used, a volume of hydrogen remains, exactly equal to what would result from the

action on water of the quantity of potassium employed. According to Thenard, the ammonia is divided into two portions; one is decomposed, so that its nitrogen combines with the potassium, forming a nitroguret of that metal, while its hydrogen is liberated, and the other portion is absorbed by the nitroguret. For further particulars respecting these strange compounds, the reader may consult the last edition of Ure's dictionary.

We are next to speak of those compounds which are known as the *salts of ammonia*.

*Carbonate of Ammonia.* This salt was formerly called *prepared ammonia*, *volatile salt*, *mild volatile alkali*, &c. There are several compounds of carbonic acid and ammoniacal gas, named in the works on chemistry, viz. the carbonate, bi-carbonate, and sesqui-carbonate, each of which shall receive some attention.

The carbonate is often called the *subcarbonate*, but its actual composition justifies the name that is now in general use; it consists, by weight, of 22 parts or one equivalent of carbonic acid, and 17 parts or one equivalent of ammonia.

The only method of forming this salt, so as to have it absolutely pure, is to mix dry carbonic acid over mercury, with twice its volume of ammoniacal gas. As soon as these gases come in contact, white clouds appear, which are soon condensed in a dry, white, volatile powder, having an ammoniacal odor, and an alkaline reaction. But it is usually made by the following process. Take a pound of muriate of ammonia, (sal ammoniac,) and a pound and a half of prepared chalk, (carbonate of lime,) previously reduced to powder. Mix them in an earthen retort, and sublime, by raising the heat until the retort becomes red.

The following diagram shows how the decomposition of the muriate of ammonia and the carbonate of lime gives rise to new products.

Before decomposition.			After decomposition.	
Muriate of ammonia	{ ammonia	17	39 carb. of ammonia.	
	{ hydrogen	1	9 water.	
	{ chlorine	36		
Carbonate of lime	{ carbonic ac.	22		
	{ oxygen	8		
	{ calcium	20	56 chloride of calcium.	
		104	104	

This result is to be regarded as the necessary consequence of chemical action, uninfluenced by any counteracting circumstance. But it often happens, that the heat is not duly regu-

lated, and then there will be some variety in the products. Thus, some carbonate and bi-carbonate will be formed; a part of the carbonate losing all its ammonia, and the carbonic acid joining a portion of the undecomposed carbonate, gives rise to some bi-carbonate, and also to the sesqui-carbonate.

The carbonate is generally met with in hard, translucent masses, of a striated appearance; it has a pungent odor, a sharp, urinous taste, and acts upon vegetable colors, in the same way as the pure alkalis. It should be kept in well stoppered bottles, for when exposed to the air it gradually loses ammonia, becomes opaque, pulverulent, and less pungent, ultimately passing into the state of bi-carbonate. Thus, suppose two equivalents of the carbonate are exposed to the air; one equivalent of ammonia will escape, and the remaining mass will consist of two equivalents of carbonic acid, and one of ammonia, which agrees, precisely, with the composition of the bi-carbonate.

One of the varieties of *smelling salts*, consists of the carbonate of ammonia, coarsely bruised and mixed with some oil of lavender, taking care to keep the mixture in well stoppered bottles.

*Bi-carbonate of ammonia*, as we have already stated, is composed of two equivalents of acid to one of alkali. It was formed by Berthollet, by passing a current of carbonic acid gas through a solution of the carbonate, as long as any of the gas appears to be absorbed. On evaporating the solution by a gentle heat, the bi-carbonate is deposited in small prismatic crystals, destitute of taste and smell.

*Sesqui-carbonate of Ammonia*. Notwithstanding all that is said of the *sesqui* compounds, I believe they will prove to be adventitious, resulting from the undue action of heat and other causes. In precise agreement with this view, is the statement of Dr Turner, that although the materials usually employed to form the sesqui-carbonate, are in proper proportion to yield the real carbonate, yet from the heat employed, part of the ammonia is lost, and the acid acquires such an ascendancy as to constitute the *sesqui* form. Indeed, from the very nature of ammonia, it is impossible to speak with definitiveness of any of its salts, usually called carbonates. Exposure to the air, and incautious application of heat, occasion an escape of ammonia, and thus lessen the relative proportion in the compound.

The incompatibles of the varieties of carbonate of ammonia are, the *acids*, *fixed alkalis* and their carbonates, *lime*, *magnesia*, *alum*, *bi-tartrate of potash*, and all the *acidulous salts*, *sulphate of magnesia*, *acetate of mercury*, *calomel* and *corrosive sublimate*, *acetate of lead*, *tartrate of iron*, and the *sulphates of iron* and *zinc*. Owing to the tendency of ammonia to escape from these com-



pounds, especially when heat is applied, they should never be added to decoctions or infusions, unless they are quite cold.

The carbonate of ammonia, if pure, is entirely volatilised by heat; if any thing remain, it may be considered as foreign matter. A large quantity of the carbonate is prepared from the soot of coal, and this is always accompanied with a fetid smell.

The carbonate of ammonia is one of the articles employed in detecting the presence of magnesia, in mineral waters; but of this, more will be said in another part of the volume. This salt is used in forming the *spiritus Mindereri*, and will be noticed under the *acetate of ammonia*. It is also employed in preparing the *cuprum ammoniatum*, the consideration of which is deferred until we examine the preparations of copper.

*Sulphate of ammonia*\* is easily prepared, by adding dilute sulphuric acid to carbonate of ammonia, until the alkali is completely neutralised. It is also contained, largely, in the soot accumulated from the burning of coal, and being quite soluble in water, is readily separated. It dissolves in two parts of water, at 60°, and in an equal weight of boiling water; on evaporating the solution, prismatic crystals are formed, composed of 40 parts or one equivalent of acid, and 17, or one equivalent of ammonia. In addition to these, the crystals contain one equivalent of water, according to Thomson, and two, agreeably to Berzelius.

From the same soot, the impure and fetid carbonate of ammonia is obtained, as we have already hinted.

*Muriate of ammonia* or the *sal ammoniac* of the shops, was originally found native, near the temple of *Jupiter Ammon*, in Africa, and hence its name. It is prepared, either by combining ammonia directly with muriatic acid, or by decomposing the sulphate of ammonia by means of muriate of soda, or the muriates of lime and magnesia by means of ammonia.

This article was first fabricated for commercial purposes, in Egypt. In that country, the dung of animals, especially of the camels so much there employed, constitutes the chief part of the fuel. On exposing the soot, formed in the burning of this fuel, to heat in glass vessels, thick masses of muriate of ammonia are sublimed on the upper part of the vessels.

We may also form it, for the purpose of a class experiment, by mixing equal measures of muriatic acid gas and ammoniacal gas, in a glass vessel, over the mercurial trough. When these gases meet, a white cloud is formed, and the condensed gases settle on the sides of the vessel.

\* Formerly known as *Glauber's secret sal ammoniac*.

When it is prepared from sulphate of ammonia, by the addition of muriate of soda, two compounds are obtained, viz. the muriate of ammonia and sulphate of soda. The dry sulphate of ammonia and muriate of soda, being mixed, are decomposed by a subliming heat. The muriate of ammonia rises to the top of the vessel, while the sulphate of soda remains at the bottom.

It is also obtained, by steeping animal matters in a strong solution of the muriate of magnesia, and exposing them to heat after they have been dried, so as to undergo a kind of slow combustion. The muriatic acid leaves the magnesia, and combines with ammonia formed by the decomposition of the animal matter. This is Mr Astley's ingenious process for the preparation of this valuable salt.

Muriate of ammonia is soluble in three parts of water at  $60^{\circ}$ , and produces cold during its solution. Boiling water dissolves its own weight of this salt, and very beautiful feathery looking crystals are obtained when it cools. It is composed of 37 parts or one equivalent of muriatic acid, and 17 parts or one equivalent of ammonia.

At a temperature below redness, this salt sublimes without fusion or decomposition, and condenses on cool surfaces, as an anhydrous salt, which attracts humidity from a moist atmosphere, but is not deliquescent, if quite pure.

The *incompatibles* of muriate of ammonia are, the *sulphuric* and *nitric* acids, both of which disengage muriatic acid gas; also, *potash*, *carbonates of soda*, *potash*, *lime*, *magnesia*, &c., whose bases detach the muriatic acid from its connexion with the ammonia. The common *sugar of lead* (acetate of lead) and the *nitrate of silver*, also decompose it.

A very useful plaster is employed in medical practice, into which the muriate of ammonia enters, and is actually decomposed. This, although an incompatible mixture, chemically speaking, is practically useful. The plaster consists of *soap* one ounce, *lead plaster* two drachms, melted together; when the mixture is cold, a half drachm of powdered sal ammoniac is added. It is easy to see, that the alkali of the soap will combine with the muriatic acid of the sal ammoniac, forming thereby a muriate of potash or soda; and during the decomposition thus effected, the ammoniacal gas of the muriate of ammonia is escaping, and on this depends the utility of the plaster. It is plain, that a renewal must be made, at least every twenty-four hours, for so soon as all the ammonia shall have escaped, the stimulant and rubefacient powers of the plaster must cease to be very obvious.

The muriate of ammonia is seldom employed in medicine,

excepting as a lotion; and then the object is to apply the cold generated by its solution, or to take advantage of the stimulus of the salt, in the treatment of indolent tumors.

We may form *freezing mixtures*, by dissolving muriate of ammonia, with other substances. If 2 parts of common salt, 1 of muriate of ammonia, and 5 of snow be mixed, the temperature will sink to  $12^{\circ}$  below zero. If 10 parts of common salt, 5 of muriate of ammonia, 5 of saltpetre, and 24 of snow be mixed, the temperature falls to  $18^{\circ}$  below zero. In these mixtures, liquefaction soon takes place, and in order thereto, an immense amount of caloric must be absorbed from the surrounding air. This caloric is said to be latent or hidden, because it is not appreciable by the thermometer. The same thing occurs, to a certain extent, when snow or ice is suspended over a fire; although it takes in a large quantity of caloric, as is obvious to the most casual observer, yet the thermometer gives no indication of such an accumulation.

If equal weights of powdered nitrate of potash and muriate of ammonia be dissolved in six or eight parts of water, a very cold solution is formed, which may be conveniently substituted for ice. In all cases, the degree of cold produced, depends upon the quantity of caloric that disappears, and this again is dependent on the quantity of solid liquefied, and the rapidity of the liquefaction.

Notwithstanding all the explanations given on the subject of freezing mixtures, there is certainly much yet unexplained, and which does not appear to admit of demonstration.

*Nitrate of Ammonia.* We merely mention this salt, in this place, because of its natural relation. Of its preparation, we have already spoken, under the article protoxyde of nitrogen, and it is not necessary to repeat the statements there made.

When this salt is heated nearly to redness, it burns with a kind of explosion; hence it was formerly called *nitrum flammans*. The crystallised nitrate is composed of 54 parts or one equivalent of nitric acid, 17 parts or one equivalent of ammonia, and 9 parts or one equivalent of water.

*Acetate of Ammonia.* This saline compound was formerly called the *spirit of Mindererus*, having been introduced by him into medical practice, in the treatment of fevers. It is made by adding acetic acid to carbonate of ammonia, until the alkali is entirely neutralised. Common vinegar is frequently used, but it gives to the compound a brownish tinge, owing to the presence of carbonaceous or cupreous matter. The colleges, therefore, recommend the use of distilled vinegar.

The acetate of ammonia does not easily crystallise. If the

liquid be placed under the receiver of an airpump, alongside of a saucer of sulphuric acid, as soon as the air is exhausted, the salt will concrete in fine, delicate crystals, which are nearly neutral. The sulphuric acid absorbs the water which escapes from the acetate by exhausting the receiver, and holds it in combination by superior affinity.

Four pints of distilled vinegar require about seven drachms of the recently prepared carbonate of ammonia, for saturation; but the strength of the distilled vinegar and the composition of the carbonate are liable to variation, so that the safest method is, to add the carbonate to the distilled vinegar, till the turmeric and litmus papers show that the neutralisation is complete.

The solution, prepared as above, is generally of a brownish tint, but it may be rendered colorless and pelucid, by filtering it through a little well burnt and fresh powdered charcoal. It will answer for medicinal use, without this purification, but should be kept in close vessels.

It has been proposed to make the *spiritus Mindereri*, in such a way as to retain all the carbonic acid gas. If the strength of the acid and carbonate be known, we can add them in the exact proportions, necessary to neutralise the alkali. The materials are then to be introduced into a proper sized vial, and the cork being fixed in its place, must be secured by the application of a piece of bladder or fine leather. The carbonic acid is displaced from its connexion with the ammonia, but is held in solution, and renders the medicine much more grateful to delicate stomachs.

The solution of the acetate of ammonia has been called *liquor of the acetate of ammonia*, *water of the acetate*, &c.

The *inecompatibles* are, the *acids*, *fixed alkalis*, *alum*, *lime water*, *sulphate of magnesia*, *corrosive sublimate*, *nitrate of silver*, and the *sulphates of zinc*, *copper*, and *iron*. If carbonic acid be present in the acetate, a precipitate will be formed, on adding a few drops of a solution of *acetate of lead*; this precipitate is the carbonate of lead. *Magnesia*, also, decomposes the acetate of ammonia, and is, of course, incompatible. It liberates a portion of ammonia, which gives pungency to the solution, and it combines with the remainder of the acetate, forming a triple salt of acetate of magnesia and ammonia.

*Phosphate of ammonia* is formed, by saturating phosphoric acid with ammonia; or by the double decomposition of bi-phosphate of lime and carbonate of ammonia. When equal parts of the phosphate of ammonia and phosphate of soda are mixed in solution, and crystallised by slow evaporation, a triple salt is obtained, called the *phosphate of soda and ammoniac*, and also

*microcosmic salt.* It may be formed, likewise, by dissolving one equivalent of muriate of ammonia, and two equivalents of phosphate of soda, in a small quantity of boiling water. As the liquid cools, prismatic crystals of the triple phosphate are deposited, while muriate of soda remains in solution. The nature of the reaction will be seen by reference to the diagram.

Before decomposition.			After decomposition.		
120 or 2 equiv. of phos. soda	{	phosph. acid	28	45 phos. amm.	60 phos. soda.
		do. do.	28		
	{	soda	32		
		soda	32		
54 or 1 equiv. of muriate of ammonia	{	ammonia	17	69 mur. of soda.	
		hydrogen	1		
		chlorine	36		
		<hr/> 174 <hr/>			<hr/> 174 <hr/>

The phosphate of soda and ammonia, according to the analysis of Mitscherlich, is composed of one equivalent of the phosphate of soda, 60, one equivalent of phosphate of ammonia, 45, and 90, or ten equivalents of water, and this corresponds with the exhibition in the diagram.

The microcosmic salt is much employed, as a flux, in experiments with the blowpipe. When heated, it parts with its water and ammonia, and a very fusible bi-phosphate of soda remains.

*Succinate of ammonia* is formed, by adding ammonia to succinic acid, until saturation is effected. It is generally employed in solution, and almost exclusively, as a test for iron. It is especially adapted to chalybeate mixtures, containing manganese. The succinic acid of the succinate combines with the peroxyde of iron, and an insoluble, ferruginous succinate is produced, of a deep red color, while the succinate of manganese is soluble.

*Benzoate of ammonia* is a compound of benzoic acid and ammonia. It is made by adding the two ingredients, until the mixture is in a perfectly neutral state. By evaporation, it may be obtained in form of feathery, but deliquescent crystals. When it is added to neutral solutions of peroxyde of iron, it gives a yellowish-red precipitate, which is a benzoate of iron. Hence, it is employed as a test for that metal.



## CHAPTER IV.

Of Lithium.

THIS substance is placed among the alkalifiable metals, but as we know very little of either the metal or its alkali, we shall say but little concerning it. In the year 1818, the substance called *lithia* was discovered by a young chemist in the employment of Berzelius, named Arfwedson. He found it in a mineral obtained from a Swedish mine, and detected in it decidedly alkaline properties. This led Sir H. Davy to examine it by the aid of galvanism, and he succeeded in demonstrating that it had a metallic base.

The minerals from which lithia is obtained, are exceedingly rare, and they do not contain more than eight per cent. of the alkali. Berzelius has since found it, in the waters of Carlsbad, in Bohemia, but the quantity of the alkali, collected from these several sources, is very small.

Lithia is a protoxyde of lithium, composed of ten, or one equivalent of lithium, and eight, or one equivalent of oxygen. It is susceptible of combination with acids, and thus forms a variety of salts, which are not of sufficient importance to call for special notice.

---

## ORDER SECOND.

## OF ALKALINE GEOFIABLE METALS.

By this term, we mean, as formerly explained, those metallic bases which, by union with oxygen, produce earthy substances, which exhibit more or less of alkaline properties, and are generally known, as the *alkaline earths*. They are four in number, viz. *calcium*, *barium*, *strontium*, *magnesium*.

---

## CHAPTER I.

Of Calcium.

THIS is the name given to a metallic substance, which, united to oxygen, constitutes lime. It has been obtained, hitherto, only in very small quantities, by decomposing carbonate of lime,

by a galvanic battery. The method of proceeding is, to make the carbonate into a paste with water, and place it on a platina tray, connected with the positive pole of the galvanic battery, putting a globule of mercury in a small cavity on its surface, and bringing it in contact with a wire from the negative pole. The lime in the carbonate is decomposed, its oxygen being disengaged, and the calcium attracted to the negative pole, where it combines with the mercury, and forms an amalgam; by exposing this to heat, the mercury is volatilised, and the calcium remains. Its equivalent is 20, and its specific gravity 2.3.

*Lime* is the most important compound of calcium, and is usually procured by exposing the carbonate of lime to a red heat, when its carbonic acid is completely expelled. For most chemical purposes, the lime that is employed for mortar will do very well; but when it is required very pure, it is obtained, most easily, by exposing white marble to heat, in a crucible placed in a furnace, for two hours; or if a larger quantity should be required, it will be found more convenient to put a good fire in the furnace, and then to fill it up with alternate layers of coals and white marble, broken into pieces not more than an inch or an inch and a half thick, and from one to three or four inches long, taking care to build them in such a manner, as may allow the fire to burn slowly and steadily. When the intervening spaces are too open, and the draught of the furnace very great, the fuel is speedily consumed, and only the external parts of the marble are decarbonated; and if a very strong fire is kept up for a considerable time, instead of obtaining the lime in solid masses of the same form as the marble put in, it generally falls into powder, which mixes with the ashes of the fuel.

Pure lime, termed also caustic lime, or quick-lime, has a fine, white color, and an earthy appearance; its taste is burning and alkaline; it corrodes animal and vegetable substances, is easily reduced to powder, and must be kept in close vessels, as it soon attracts water and carbonic acid from the air, crumbling down into a fine powder. It phosphoresces powerfully when heated to full redness. It is one of the most infusible bodies known; fusing with difficulty, even by the heat of the oxy-hydrogen blowpipe. A small piece of fresh lime, subjected to the action of the mixed gases, yields a light equal to that of twenty spermaceti candles.

When water is poured upon lime, it combines with it immediately, forming *hydrate of lime*, and great heat is evolved during the combination, arising from the water giving out its caloric and entering into the solid state. The compound is quite solid, though in the form of a bulky powder, and is usually termed

*slaked lime* ; and as it absorbs carbonic acid by exposure to the air, it should be kept in close bottles. The hydrate of lime differs from the hydrates of barytes and strontian, by parting with its water, at a red heat.

The equivalent of lime is 28, being composed of calcium 20, or one equivalent, and oxygen 8, or one equivalent. Of the use of lime, in detaching carbonic acid from potash and thus rendering it more caustic and solid, we have already spoken. It is also important, as the basis of many lutes, and of the common mortar used in building, and also of cements in general. Although the lime in ordinary mortar, is nearly free of carbonic acid, yet by long standing it absorbs its original dose, so that the mortar in very old buildings, is often found as compact as limestone. This change is partly, however, attributable to a crystalline arrangement, connected with the solidification of the water absorbed with the carbonic acid.

Lime is also used by the tanner, to detach the hair from hides; and in other departments of the arts, it is of very great importance.

Lime is sparingly dissolved by water, yet in sufficient quantity to impart its characteristic properties. The solution, which is known by the name of *lime water*, is made by agitating slaked lime, repeatedly, in a bottle containing pure water. At the end of twenty-four hours, the clear liquid may be decanted and put in close vessels, for use. Or a piece of pure lime, just from the kiln, may be thrown into a stone jar, to be slaked; after which the requisite quantity of water may be poured in, and the whole well stirred for the space of fifteen minutes. Then let it settle, and pour off the clear liquor, as directed above. A piece of lime, of the size of a hen's egg, will be amply sufficient for ten gallons of water.

The *milk* or *cream* of lime is lime water, in which some hydrate of lime is suspended. It has, of course, a thicker consistence than lime water. Lime water (the aqua calcis of the shops) turns the infusion of blue cabbage to a green, and turmeric to a brown. It gives a precipitate of carbonate of lime, with carbonic acid water; and by breathing through it with a bent tube, the same kind of precipitate is formed, if an excess of carbonic acid be not added. By continuing to breathe into the fluid for some minutes after it has exhibited the turbid or milky appearance, it will be found to regain its original transparency, owing to the formation of a soluble bi-carbonate of lime. Exp.

Lime water gives no precipitate with sulphuric acid, because sulphate of lime is more soluble than lime itself, although only very sparingly soluble in pure water. Exp.

If some lime water be exposed to a boiling heat in a Florence flask, part of the lime will be precipitated, showing that it is more soluble in cold, than in hot water. According to Mr Dalton, one part of lime requires upwards of 700 parts of cold water for solution, and boiling water dissolves little more than half that quantity.

The most delicate test of lime is oxalic acid. If a few drops of a solution of this acid be added to the most dilute solution of lime that can be made, we shall instantly perceive a milky precipitate of oxalate of lime, that being less soluble in water than any other salt of lime.

A few simple experiments may be easily made by any student, to render the nature of lime, its incompatibles, and chemical relations perfectly familiar.

Pour some sulphuric acid on a little quick-lime placed in an iron cup, taking care to avoid the fumes; sulphate of lime is formed, and the heat produced by the combination is so great, that part of the sulphuric acid is volatilised. If a little gunpowder is thrown upon it, immediately after the sulphuric acid has been added, it is inflamed.

Mix some dilute nitric or muriatic acid, with a little carbonate of lime; the latter will be decomposed, its lime uniting with the muriatic acid, and forming a soluble muriate of lime, while the carbonic acid is disengaged with effervescence.

Add a little sulphuric acid to about an equal bulk of a saturated solution of muriate of lime; sulphate of lime is formed, copious fumes of muriatic acid are disengaged, and a solid mass of sulphate of lime remains. Dilute the acid and the solution with equal bulks of water, previously, and mix them together, after allowing the acid to cool; the mixture becomes nearly solid, sulphate of lime being formed as before, and the water mixed with it retaining almost all the muriatic acid in solution. If, however, the solution is diluted with about 500 or 600 parts of water, then no precipitation takes place, the sulphate of lime being soluble even in a smaller quantity of water.

Add a solution of potash, or soda, to a solution of any soluble salt of lime; the alkali unites with the acid, and lime is precipitated.

Add a little carbonate of potash, or soda, to lime water, or a salt of lime in solution; carbonate of lime is precipitated, and when the lime has been combined with an acid, the potash or soda unites with it, as in the preceding experiment, and remains in solution.

Drop a small quantity of oxalic acid into lime water; a copious precipitate of the oxalate of lime immediately appears. Add a little oxalic acid to a solution of the sulphate of lime; a slight precipitate only appears, and if an excess of sulphuric acid should have been previously mixed with it, no apparent change takes place. The reason is, that sulphuric acid has a stronger affinity for lime than oxalic acid; if, however, ammonia, soda, or potash, be added to neutralise the acid, then the oxalic acid combines at once with the lime, and a copious precipitate appears. Hence, oxalate of ammonia or oxalate of potash are employed, in preference to pure oxalic acid, for precipitating lime from its solutions, and when an excess of acid is present, it must be neutralised before we begin to apply our tests. Exp.

Lime water is much employed in medical practice, internally and externally. Its *incompatibles* are, *all alkaline and metallic salts, borates, tartrates, citrates, sulphur, spirituous preparations, and the infusions of all astringent vegetables.*

Lime, in fine powder, is occasionally blown into the eyes of workmen, especially the laborers attendant on house-builders. If the lime be fresh, its caustic action is extremely painful, and unless neutralised or blunted, will soon excite extensive inflammation. The late Professor Wistar, (University of Pennsylvania,) was in the habit of injecting sweet oil into the eyes in such cases, with complete success. The oil and lime formed a mild liniment, which alleviated the symptoms, on the same principle on which it acts in mitigating the pain of an ordinary burn.

Lime may be confounded with barytes and strontian. The surest distinguishing marks are the following; nitrate of lime yields prismatic crystals by evaporation, which are deliquescent in a high degree, and quite soluble in alcohol. The nitrates of barytes and strontian also crystallise, but the crystals do not deliquesce, unless the air be very moist, and they do not dissolve in pure alcohol.

The salts of lime, when heated before the blowpipe, or when their solutions in alcohol are set on fire, communicate to the flame a dull, brownish-red color.

In addition to the *protoxyde* of calcium, (lime,) there is a *deutoxyde* mentioned in some of the books, which contains one equivalent of calcium and two of oxygen. Calcium combines with chlorine, sulphur, and phosphorus, forming chloride, sulphuret, and phosphuret of calcium; but as the two first will be best studied after the examination of the muriate and sulphate of lime, they shall be deferred for the present.



*Phosphuret of calcium* is formed, whenever phosphorus is brought into contact with lime, at a high temperature. An easy method is, to throw some pieces of dry phosphorus into a crucible, with a few fragments of fresh lime, (about the size of a pea). The crucible should be red hot, and inverted in a sand bath, as soon as the phosphorus and lime have been put in. Sand should be heaped round the edge of the crucible, to prevent the escape of any part of its contents. The experiment may be performed with 20 or 30 grains of phosphorus, and 60 or 70 of lime. It may also be made, by taking two crucibles, one three times as large as the other, and heating 60 or 70 grains of lime to redness in the larger crucible. The smaller vessel is placed in a sand bath, and should contain 20 or 30 grains of phosphorus. The larger crucible should be speedily removed from the fire, and its contents cautiously but quickly poured upon the phosphorus; and then it should be inverted over the small crucible, sand being heaped around its edge, to make it air-tight, or nearly so. I have employed this method, very successfully, and think it the easiest and most certain. When the crucibles are cold, the contents should be placed in well stoppered bottles.

Some writers are of opinion, that the phosphuret of calcium is a compound of phosphate of lime and phosphuret of calcium; while others call it by the original name, viz. *phosphuret of lime*. If a phosphuret of calcium be formed, the lime must be decomposed, and while its calcium joins a portion of the phosphorus, it is probable that its oxygen unites with another portion, to form phosphoric acid. This latter may then unite with some undecomposed lime, and phosphate of lime may result. That such may be the fact, will be seen from the following diagram.

36 or 3 equiv. of phosphorus	{ phosphorus 12			
	{ phosphorus 12			
	{ phosphorus 12			
28 or 1 equiv. of lime	{ calcium 20		32	{ phosphuret of calcium.
	{ oxygen 8			
28 or 1 equiv. of lime	{ calcium 20		32	{ phosphuret of calcium.
	{ oxygen 8			
28 or 1 equiv. of lime	28		56	{ phosphate of lime.
<hr/> 120 <hr/>	<hr/> 120 <hr/>		<hr/> 120 <hr/>	

The properties of phosphuret of calcium were fully set forth, when phosphuretted hydrogen gas was under consideration.

*Muriate of lime*, called, also, hydrochlorate of lime, is prepared, by dissolving carbonate of lime in dilute muriatic acid. The muriatic acid combines with the lime, and carbonic acid gas flies off rapidly; enough acid should be used to saturate the lime completely, or to form a neutral muriate. This salt is crystallised with some difficulty, on account of its tendency to absorb moisture from the atmosphere. It is very soluble, both in water and alcohol, and when exposed to heat, so as to lose seven equivalents of water, a chloride of calcium remains. If water be added to the chloride thus formed, it returns to the state of muriate, precisely in the manner illustrated, when treating of chloride of sodium and muriate of soda. The following diagram shows the conversion of muriate of lime into chloride. The water that is also formed, is driven off by heat.

Before decomposition.		After decomposition.	
37 muriatic acid	{ hydrogen chlorine	1 36	9 water.
28 lime	{ oxygen calcium	8 20	56 chloride of calcium.
—	—	—	—
65 muriate of lime.	65	65	
—	—	—	—

The 65 parts of muriate of lime, in the diagram, are exclusive of the 45 parts of water, found in the salt; but these, as well as the nine parts formed by decomposition, escape under the influence of proper temperature.

The exact constitution of crystallised muriate of lime is, 37 muriatic acid, 28 lime, and 45 water; the whole equivalent being 110.

The crystallised muriate is the compound which produces such an intense degree of cold, when mixed with snow. It is prepared for this purpose, by evaporating the solution until a drop of it becomes solid so soon as it falls on a cold saucer. Three parts of this article, mixed with two of freshly fallen snow, cause the mercury in the thermometer to fall  $50^{\circ}$  below the freezing point. The strong affinity of this salt for water, causes it to liquefy rapidly, and in becoming fluid, it absorbs an immense quantity of caloric; and just in proportion as this becomes latent, is the intensity of the cold resulting from the operation.

In addition to the above method for procuring the *chloride of calcium*, it may be formed by passing chlorine gas over pure lime, exposed to a red heat, in a porcelain tube. The chlorine decomposes the lime, unites with its calcium, and oxygen escapes.

This chloride deliquesces rapidly, on exposure to the air; it has a very bitter and disagreeable taste, and is dissolved by about one-fourth its weight of cold water. It is also soluble in alcohol, and imparts a red color to its flame, when it is kindled.

The crystals of the chloride are employed to form freezing mixtures, in the same manner as the muriate.

Both the crystallised muriate and the chloride are employed to free alcohol of water. They absorb this fluid entirely, and the alcohol can then be separated by distillation.

What is called the *oil of lime*, is nothing more than the dry muriate, changed by exposure to the air.

*Homborg's phosphorus* is formed, by exposing the muriate of lime to a heat sufficient to condense it so completely, that, on cooling, it appears to be a vitreous mass. On striking this mass with a hard substance, it emits a phosphoric light.

The solution of muriate of lime, is used in medicine, chiefly in scrophulous affections. It is made, by dissolving two ounces of the muriate in three ounces of distilled water.

*Sulphate of lime* is easily prepared, by mixing a solution of muriate of lime, with any soluble sulphate; or by adding diluted sulphuric acid to powdered chalk. It is obtained, very largely, as a residuum in the manufacture of the artificial Seltzer water.

This salt occurs very extensively, as a natural product. The mineral called *anhydrite*, is anhydrous sulphate of lime; and all the varieties of *gypsum*, or *selenite*, or *plaster of Paris*, are composed of sulphuric acid and lime. The whiter and more compact species of the sulphate of lime, are used by the statuary, under the name of *alabaster*, a term which seems to have, among artists, at least, no very definite meaning.

Pure sulphate of lime is composed of 40 or one equivalent of sulphuric acid, and 28 or one equivalent of lime; in the crystallised state, there are 18 additional parts or two equivalents of water.

The hydrous salt is deprived of its water of crystallisation, by a low, red heat, and in this state, constitutes the plaster of Paris. Its property of becoming hard, when made into a thin paste with water, is owing to the anhydrous sulphate combining chemically with that liquid, and thus depriving it of its fluidity.

The plaster is then said to set, or to be fixed, and on account of this property, it is highly recommended in a late number of the London Mechanics' Magazine, as a safe and certain poison for rats. The plaster is mixed with an equal quantity of oat-meal, and on water being taken into the stomach, the plaster becomes hard, and proves fatal to life. It is affirmed, that no animal but the rat will eat this dry mess.

Sulphate of lime is nearly void of taste, and has no smell. It is partially soluble in water, requiring 500 parts of cold, and 450 of boiling water, for its solution. Owing to this circumstance, and to its general prevalence in the earth, it frequently enters into the composition of spring water, giving it the property called hardness.

It has been affirmed, that this salt will bear a white heat, without experiencing decomposition, but Dr Thomson says, it loses a part of its acid at a red heat. Be this as it may, we are sure, that when mixed with charcoal in a crucible, and exposed to a red heat, it is completely decomposed, and *sulphuret of calcium* is obtained and carbonic oxyde is disengaged. The same play of affinities occurs in this process, as in the formation of sulphuret of potassium, by decomposing sulphate of potash, and we refer our readers to the diagram on that subject. The sulphuret of calcium is often called sulphuret of lime, but they are different articles. By acting on the sulphuret of calcium, by means of water, we obtain the *hydrosulphuret of lime*, and the sulphuretted hydrosulphuret may be prepared, by a process analogous to that already given for making the sulphuretted hydrosulphuret of potash.

*Carbonate of lime* is an abundant natural product, being the basis of marble, chalk, limestone, and a variety of other substances. We can form it, as we have already shown, by precipitation; but although insoluble as a carbonate, it is easily dissolved when carbonic acid is in excess. The spring water of limestone districts holds the bi-carbonate of lime in solution, and the reason why lime is deposited in the kettles, on boiling, is because the excess of carbonic acid is driven off by the heat, and an insoluble carbonate remains.

The *prepared chalk*, *prepared oyster shells*, *crabs' claws*, &c., of the shops, are all varieties of the carbonate of lime. When these are administered, as medicinal agents, it is generally with a view to chemical effect; hence they are frequently given as antacids. The acid of the stomach, whatever it may happen to be, combines with the lime, and the carbonic acid of the carbonate escapes, by eructation.

*Nitrate of Lime.* We may prepare this salt by adding dilute nitric acid to powdered chalk, until effervescence ceases. If the solution be evaporated, prismatic crystals will be formed, which are deliquescent, and soluble in alcohol, as well as in water. It contains one equivalent of each of its component parts.

This salt is the *calcareous nitre* or *nitrous selenite*, from which, by the addition of potash, saltpetre is obtained.

Exp. If nitrate of lime be fused in a crucible, and then exposed for some time to heat, it becomes phosphorescent, forming *Baldwin's phosphorus*.

Exp. If nitrate of lime be dissolved in alcohol, it imparts a red color to the flame.

*Fluate of Lime.* This is the basis of the mineral called *fluor* or *Derbyshire spar*, which is exceedingly abundant, as a natural product. We noticed the different views respecting the actual composition of the spar, when treating of fluoric acid. By some, it is regarded as a compound of fluoric acid and lime, while others say, that it consists of fluorine and calcium, and hence they call it the fluoride of calcium.

The *chloride of lime* is the last of the compounds of lime, to be noticed in this place. It is the article usually sold under the name of *bleaching salts*, and was formerly called *oxymuriate of lime*. It is materially different from chloride of calcium, already examined. It is prepared on the large scale, by exposing slaked lime to the action of chlorine gas, which soon combines with it, with evolution of heat, which is very apparent, when the absorption of the gas proceeds rapidly. A solution, which may be used for all the purposes to which this compound is usually applied, may be prepared, by transmitting a current of chlorine, (prepared from an ounce or two of the peroxyde of manganese with the proper quantity of muriatic acid,) through two ounces of the hydrate of lime, suspended in twelve or fourteen ounces of water. A similar solution may be obtained, by dissolving the chloride of lime, of commerce, in water.

When the dry chloride is exposed to heat, a considerable quantity of chlorine is disengaged at first, which is succeeded by oxygen gas, a portion of chloride of calcium then being formed. The same change takes place, when it is kept for a long time, even in the dark, and in close vessels; so that on opening a bottle full of this compound, which has been accurately closed a considerable time before, a quantity of oxygen immediately escapes; and often throws some of the chloride out of the bottle, a circumstance that the experimenter must always bear in mind to prevent any of it being thrown in his eye; an accident that frequently happens on taking out the cork inadvertently.

The solution of the chloride of lime is applied to the same valuable purposes, as the solution of chlorine in water; acids decompose it, combining with the lime and discharging the chlorine. To obtain the full benefit of the chlorine in disengaging vegetable colors, an acid must be added to it to combine with the lime, and set the chlorine at liberty.

The value of chloride of lime depends on the quantity of chlo-



rine it contains, and this differs considerably, not only according to the care used in preparing it, but also the precautions taken to preserve it; and hence, several methods for ascertaining its strength have been proposed. One consists, in adding a given quantity of the diluted solution to a solution of indigo in sulphuric acid, of a known strength; the strength of the chlorine being indicated by the quantity of the solution which it can decolorise. M. Morin has proposed a solution of the muriate of manganese as a substitute for this, as he affirms it is more to be depended on, in giving accurate indications of the strength of the chloride; the lime combining with the muriatic acid, and precipitating the brown oxyde of manganese, while the chlorine is disengaged. The quantity of the muriate decomposed, corresponds exactly with the chlorine set at liberty.—*Quarterly Journal, New Series*, vol. vi.

The actual composition of chloride of lime seems to be a matter of uncertainty, and from the very nature of chlorine, especially its tendency to escape, we should suppose it difficult to decide upon any actual standard, excepting that connected with its decolorising power. Dr Ure analysed three specimens of chloride, and in 100 parts of each, the proportions of chlorine were, as 23, 40.32, and 39.5. The generally received composition is, 36, or one equivalent of chlorine, and 28, or one equivalent of lime. The remarks made respecting the union of chlorine with the oxyde of sodium, are equally applicable to its relation in the present case.

---

## CHAPTER II.

### Of Barium and Strontium.

As these metals are formed in the same manner, and have similar properties, we have placed them in one class.

It is proper, however, to notice, that a record is made in the Elements of Lavoisier, of the reduction of the metallic base of barytes, and the name given to the metal, by the operator, was *barytum*.\*

Sir H. Davy obtained barium and strontium, by a process precisely similar to that which has been described for making calcium.

\* See page 223.

Barium is of a dark-gray color, with a lustre inferior to that of cast iron. As it sinks in strong sulphuric acid, its specific gravity must be greater than that of the acid itself. It attracts oxygen rapidly from the air, and is thus changed into a white powder, which is barytes, or the protoxyde of barium. When barium is thrown into water, that fluid is decomposed, the oxygen uniting with the barium, while the hydrogen escapes with effervescence. The chemical equivalent of this metal is 70.

*Oxydes of Barium.* The well known barytes is the first oxyde of this metal, to be noticed. It was discovered by Scheele, in 1774, and was called by him, from a Greek word, signifying heavy, because it was the heaviest earth, then known.

Barytes may be prepared from the nitrate, by exposure to a red heat; or from the carbonate, by a white heat. It should be reduced to powder, and intimately mixed with charcoal, also in powder, and then exposed to heat in a black lead crucible.

Barytes is a grayish-white powder, the specific gravity of which is about 4. It requires a very high temperature for fusion, but melts before the oxy-hydrogen blast. It has a sharp, caustic, alkaline taste, changes vegetable blues to green, and neutralises the strongest acids. Its alkaline character is, therefore, as decisive as that of potash or soda, but it is less caustic and less soluble in water than those alkalis. In pure alcohol, it is quite insoluble. Its affinity for water is very strong, and when mixed with that liquid, it slakes as quick-lime does; the heat evolved in the slaking is so intense as to be, sometimes, luminous. The result of this union is a white hydrate of barytes, which is fusible at a red heat, but which retains its water, at the highest heat of a smith's forge. It is composed of 70 parts or one equivalent of barium, 8 parts or one equivalent of oxygen, and 9 parts or one equivalent of water.

The hydrate of barytes dissolves in twice its weight of water at 60°, and the solution is an excellent test of the presence of carbonic acid, in the atmosphere or in other gaseous mixtures. The carbonic acid unites with the barytes, and a white, insoluble carbonate subsides.

Barytes is distinguished from all other substances, by the following marks. Its soluble salts give white precipitates of carbonate of barytes, when alkaline carbonates are added; and sulphate of barytes, which is insoluble in acid and alkaline solutions, when sulphuric acid or any soluble sulphate is added. It forms, in union with muriatic acid, a salt that is insoluble in alcohol, is unchanged by exposure to the air, and readily crystallises after having been duly evaporated.

The most convenient mode of forming the salts of barytes, is

by the action of diluted acids on the native or artificial carbonate. All the soluble salts of barytes are poisonous. The carbonate, although called insoluble, is acted upon by the fluids of the stomach to such an extent, as to evince poisonous properties. The sulphate, being absolutely insoluble, is wholly inert. Hence, the treatment of a case of poisoning by barytes, so far as chemical agency is concerned, consists in changing the poison to an absolutely insoluble mass, and this is best done by giving freely of the sulphate of magnesia, or Epsom salts. Part of the sulphuric acid of this salt joins the barytes and renders it inert, while the undecomposed part of the Epsom salts carries the offending matter out of the system, by its cathartic action.

The *deutoxyde* or *peroxyde of barium* (for there are but two oxides) is prepared by passing oxygen gas over barytes, at a red heat. M. Quesneville, junior, has recommended an easier process, which is as follows. Nitrate of barytes is introduced into a luted retort of porcelain, to which is attached one of Welter's safety tubes, terminating under an inverted jar full of water. Heat is applied to the retort, gradually, until a red heat is gained, which must be continued as long as any nitric oxide or nitrogen gas comes over. When these no longer appear, and pure oxygen is evolved, it is known that all the nitrate is decomposed, and the process must be discontinued. In both cases, the barytes acquires an additional equivalent of oxygen, for the deutoxyde consists of one equivalent of barium and two of oxygen.

It is this deutoxyde that was employed by Thenard, in the formation of the deutoxyde of hydrogen. When thrown into water containing a small quantity of muriatic acid, one equivalent of oxygen is set at liberty, by which the deutoxyde is changed into the protoxyde, and combines with the acid. When the process is conducted with due care, the oxygen does not escape, but unites with the hydrogen of the water, and raises it to the maximum of oxydation.

*Sulphate of barytes* or *heavy spar* exists in considerable quantity in nature, and is formed artificially, when sulphuric acid and barytes are presented to each other in solution. It is very insoluble in water, but is rendered chemically useful by being converted, in the first place, into *sulphuret of barium*, by decomposing it with charcoal, (in the same manner as in the preparation of sulphuret of potassium,) and then into hydrosulphuret of barytes, by digesting it in water. The native sulphate should be exposed to a full red heat and quenched in cold water, that it may be easily reduced to a fine powder; it is then mixed with a fifth part of its weight of powdered charcoal, and made into

a tough mass, with a little oil, after which it is exposed to a strong red heat for two hours, in a furnace. The mass that remains, consists almost entirely of sulphuret of barium, with a little undecomposed sulphate of barytes. By digestion in water, the sulphuret of barium is dissolved and a solution of the *hydrosulphuret of barytes* obtained, the sulphuret reacting on part of the water in the manner already described. If a strong solution of the hydrosulphuret is made, crystals will be deposited as it cools; it is very soluble in water.

*Nitrate of barytes* is prepared, by adding nitric acid to a solution of the hydrosulphuret of barytes till it ceases to produce any effervescence, or till the solution reddens the vegetable blues, evaporating it, after it has been filtered, till a pellicle appears on its surface. The sulphuretted hydrogen is disengaged, and a considerable portion of it is decomposed by the nitric acid, a copious precipitation of sulphur taking place. By digesting the native or artificial carbonate of barytes in nitric acid, diluted with 8 or 10 parts of water, the same compound is obtained, and carbonic acid is disengaged with effervescence. This salt crystallises in octohedrons which contain no water of crystallisation, and are soluble in about 11 parts of cold and 3 of boiling water.

*Carbonate of barytes* is very insoluble, like the other earthy carbonates, and effervesces with the stronger acids; it exists native, and when required in a state of minute division, it is usually prepared, by adding a solution of an alkaline carbonate to a solution of the hydrosulphuret, or of any soluble salt of barytes. The carbonic acid of the carbonate unites with the barytes, while the sulphuretted hydrogen, previously combined with the barytes, joins the alkali, and the compound thus formed remains in solution, while the carbonate of barytes is precipitated.

This salt is found, almost always, in lead mines, where it is called *Witherite*, after Dr Withering, its discoverer.

*Muriate of barytes* may be prepared in the same manner as the nitrate, only changing the acid. When the carbonate of barytes is employed, the muriatic acid must be diluted with thrice its bulk of water, and the carbonate, in fine powder, added as long as effervescence continues, in order to have a perfectly neutral solution. When evaporated to a pellicle, and set aside to crystallise, delicate tabular crystals are formed, composed of one equivalent of muriate of barytes and one of water. They are soluble in five parts of water at 60°, but insoluble in alcohol, and when exposed to a red heat, are changed into *chloride of barium*.

The watery solution of the muriate of barytes is much employed as a test of sulphuric acid, either free or combined. If a few drops be added to water containing a little sulphuric acid, or some sulphate of magnesia, a white precipitate of sulphate of barytes will be thrown down. Exp.

The same solution is also employed in medical practice in the treatment of scrophulous affections; but the dose is necessarily small, as the article is highly poisonous. When a large portion has been swallowed, through mistake, its effects should be prevented by the speedy and free use of a solution of sulphate of magnesia.

*Chlorate of barytes* is composed of 76, or one equivalent of chloric acid, and 78, or one equivalent of barytes. Its importance depends, chiefly, on its agency in forming the chloric acid. It is made by digesting, for a few minutes, a strong solution of chlorate of potash with a slight excess of fluosilicic acid. The potash is precipitated in the form of an insoluble fluosilicate, while chloric acid remains in solution. After filtration, the liquid is neutralised with carbonate of barytes.

Ure directs the chlorate to be made, by passing a stream of chlorine gas through a hot solution of barytes, or through hot water, in which this earth is kept suspended. Here the action is precisely similar to that which occurs in preparing the chlorate of potash.

*Strontia* or *strontites* is a compound very analogous to baryta or barytes. It is composed of 44, or one equivalent of strontium, and 8, or one equivalent of oxygen, and it is represented by the number 52. Hope and Klaproth discovered it in 1792, at about the same time, in a mineral found at Strontian, in Scotland, and called *strontianite*.

Pure barytes and strontites agree in most of their properties; and, in general appearance, they are very much alike. They may be distinguished from each other, however, by their action on inflammable matter, in a state of combustion. If we place a small crystal of muriate or nitrate of strontites on the wick of a burning candle, or if the wick be touched with a glass rod wet with a solution of a strontian salt, the flame will instantly become red. Barytes and its salts give only a faint yellow, with a tinge of green, to the flame, so that its action is scarcely perceptible.

The mixture for the red light used in fireworks, is composed, chiefly, of nitrate of strontites, with about an eighth part of sulphur and a tenth of charcoal; a small quantity of sulphuret of antimony and chlorate of potash are mixed with it, occasionally, to render it more inflammable.



The salts of strontites are made, pretty much as the salts of barytes. There is between them, however, a very important difference; the salts of strontites are destitute of poisonous properties, while those of barytes are generally deleterious.

The soluble salts of strontites are used as tests, in the same way that we apply the salts of barytes.

---

## CHAPTER III.

### Of Magnesium.

A BRIGHT metallic regulus was obtained from magnesia, at Chemnitz, in Lower Hungary, by Messrs Tondi and Ruprecht, about the year 1790, to which the name of magnesium was given. Since then, Sir H. Davy obtained this metallic base, by galvanic agency, by the same process which yielded barium and calcium. This metal decomposes water, unites with its oxygen, and is changed into magnesia. Its equivalent is 12, while that of magnesia, the only known oxyde, is 20. Its specific gravity is 2.3.

To prepare magnesia, the carbonate is exposed to a red heat in a covered crucible, for an hour or two; when a small quantity is employed, as half an ounce or an ounce, the carbonic acid is completely expelled in half an hour, if the temperature is not allowed to decline. It must be kept in well stoppered bottles, or it will soon absorb a sufficient quantity of carbonic acid to destroy its original character.

Magnesia may be obtained, also, by adding a solution of caustic potash or soda to a solution of any of its salts, the alkali uniting with the acid of the magnesian salt, and forming a compound which remains in solution, while the magnesia is precipitated. It must then be washed on a filter, by repeated affusions of water.

Throw a small quantity of the magnesia into a glass containing sulphuric acid, diluted with three times its bulk of water. If it contains no carbonic acid, it will be completely dissolved without effervescence, on stirring the liquid with a glass rod.

Mix another portion with an ounce or two of the blue infusion of cabbage; it will be rendered green in the same manner as by a solution of an alkali or alkaline earth; but water with which magnesia has been shaken does not produce

this effect, because very little is dissolved. Dr Fyfe says, that 5000 parts of water are requisite to dissolve one of magnesia, at  $60^{\circ}$ . Some persons do not regard it as an alkaline earth, but it has, undoubtedly, alkaline properties, though in a feeble degree. Pure magnesia is a white, friable powder, of an earthy appearance, and communicating a sensation of harshness to the fingers; it has neither taste nor odor. It is exceedingly infusible, has a weaker affinity for water than lime has, and yields all its water at a red heat.

Magnesia is characterised by the following properties. With nitric and muriatic acids, it forms salts which are soluble in alcohol, and very deliquescent. The sulphate of magnesia is very soluble in water, a circumstance that distinguishes it from the sulphates of lime, barytes, and potash. It is precipitated from its salts as a bulky hydrate, by the pure alkalis. It is precipitated as carbonate of magnesia, by the carbonates of potash and soda. Its salts give no precipitate, when in solution, with the carbonate of ammonia; but on adding phosphoric acid or phosphate of soda, a copious white precipitate will be thrown down, composed of phosphoric acid, ammonia, and magnesia.

Notwithstanding the great insolubility of magnesia in water, it has the property of increasing the solubility of camphor, opium, and resins, in the same fluid; it is soluble in solutions of the alkaline carbonates, but not in those of the caustic alkalis.

The names *magnesia usta* and *magnesia calcinata*, were given to pure magnesia, simply because it is made by the aid of a burning or calcining heat.

If magnesia, as we usually find it in the shops, be mixed with calomel, and the mixture stirred in a wine glass half full of water, in a few minutes an ash-gray precipitate will be seen, which I regard as the protoxyde of mercury. I <sup>Exp.</sup> have not seen this fact noticed in any of the books, but it is not difficult to give a clear view of the action that ensues on making the mixture, so as to justify this conclusion. Thus:

20 or 1 equiv. of magnesia	20	57 muriate of mag.
9 or 1 equiv. of { oxygen	8	
water { hydrogen	1	
236 or 1 equiv. of { chlorine	36	
calomel { mercury	200	208 protox. merc.
<hr/>	<hr/>	<hr/>
265	265	265
<hr/>	<hr/>	<hr/>

Having frequently observed this change in making the watery mixture of calomel and calcined magnesia, I have since pre-

ferred to mix them in syrup. The decomposition, if it occur at all in the latter, is slow in its operation, and cannot be seen by the patient.

Magnesia is preferable to the carbonate, in all cases of excessive flatulence; since the evolution of carbonic acid from the latter, would necessarily augment the difficulty. If magnesia be given as an antacid, it will neutralise the acid and then act as a cathartic; but the records of medicine furnish well attested cases of falsely supposed acidity, in which magnesia was used for months and even years, until, at length, it accumulated in large masses, in various parts of the intestinal canal.

We test the absolute purity of magnesia by acids, with which it will not effervesce, if quite pure. If magnesia and muriatic acid be placed at one time, in separate cups, in the scale of a good balance, no diminution of weight should occur, on mixing them.

To detect lime, we add to a solution of the suspected magnesia, in sulphuric acid, some oxalate of ammonia; if lime be present, there will be an insoluble precipitate of oxalate of lime.

Magnesia was, originally, a *general* term, expressive of any substance which had the power of attracting some principle from the air, from *magnes*, the loadstone. What we now call *magnesia*, was sold as a *panacea*, by a canon at Rome, in the seventeenth century, under the title of *magnesia alba*, or Count Palma's powder.

*Muriate of Magnesia.* This salt is composed of 37 parts muriatic acid, and 20 of magnesia; and, in the form of crystals, it contains 45 parts of water. It exists in many mineral springs, and is contained abundantly in sea-water. When muriate of soda is separated from sea-water, by crystallisation, an uncrystallisable fluid, called *bittern*, is left, which consists, chiefly, of muriate of magnesia, and is much employed in the manufacture of sal ammoniac, by the decomposition of sulphate of ammonia.

Exp. If muriatic acid be saturated with magnesia, the resulting compound will be muriate of magnesia.

This salt has a bitter taste, is highly soluble in alcohol and water, and deliquesces with rapidity in the open air. When heated to redness, it loses a portion of acid as well as water.

*Sulphate of magnesia* or *Epsom salts*, may be prepared, by adding carbonate of magnesia to sulphuric acid diluted with four parts of water, till it is completely neutralised, evaporating the solution afterwards till crystals are formed, when a drop of the liquid is taken out on a glass rod; it may then be set aside to crystallise. It is soluble in its own weight of cold water, and in three-fourths of its weight of boiling water.

On the large scale, it is prepared from bittern, the liquid that remains after most of the common salt has been obtained from sea-water by evaporation; it is boiled, to separate an additional portion of common salt and part of the water, crystals of sulphate of magnesia being deposited as it cools, which are purified by repeated crystallisation. It is also formed by the action of sulphuric acid on magnesian limestone, (a compound of carbonic acid, magnesia, and lime,) sulphates of magnesia and lime being produced, by the sulphuric acid combining with the magnesia and lime; the sulphate of magnesia is separated by solution in water.

The sulphate of magnesia, of commerce, is sometimes mixed with a little muriate of magnesia, derived from the bittern from which it is prepared. It renders the salt deliquescent, and may be easily detected by pouring sulphuric acid on it, and holding over it a rod dipped in water of ammonia, muriatic acid being disengaged by the sulphuric acid, if any muriate is mixed with the salt, and forming white fumes with the ammoniacal gas arising from the water of ammonia.

This sulphate is composed of one equivalent of sulphuric acid and one of magnesia; in its crystallised state, the composition is 40 sulphuric acid, 20 magnesia, and 63, or seven equivalents of water.

This salt has a nauseous and *bitter* taste; hence one of its names is, the *purging bitter salt*. It is usually seen in fine, needle-like crystals, closely resembling the crystals of oxalic acid; when it is quite pure, these crystals effloresce, on exposure to the open air.

Sulphate of magnesia, although soluble in water, is insoluble in alcohol. One fluid ounce of water, dissolves an ounce, by weight, of the salt. The unpleasant taste of the solution, may be almost entirely removed, by the addition of a small quantity of lemon juice, or cream of tartar.

The *incompatibles* of this salt are, the *muriates of ammonia, barytes, and lime; nitrate of silver; subacetate and acetate of lead. Phosphate of soda* occasions no immediate precipitate, unless ammonia be present; in that case, the triple compound will be formed, called the *ammoniaco-magnesian phosphate*. Paris says, that ammonia will effect a decomposition of the sulphate of magnesia, and Brande makes an opposite affirmation. I have tried the experiment, and know that the statement of Brande is correct.

Mr West, of Lymington, in England, has succeeded in making the sulphate in large and beautiful crystals, which cannot be mistaken for those of oxalic acid, but I have never seen a sample of them.



In order to distinguish between oxalic acid and Epsom salts, it should be borne in mind, that the acid is intensely sour, while the sulphate is bitter. If a portion of each be dissolved in water, and lime water be added, it will make the oxalic solution very milky, while it will scarcely have any effect on the other solution.

To a solution of Epsom salts, add a solution of carbonate of potash, and carbonate of magnesia will be precipitated. But if the latter solution be added to a solution of oxalic acid, there will be no precipitation, because the oxalate of potash is soluble in water.

As these substances have actually been mistaken for each other, it is important to have them distinctly and fully labelled with their plain titles.

The *phosphate of ammonia* and *magnesia*, called also the *triple phosphate*, has been noticed. It may be made by the direct union of the two phosphates, or by evaporating urine. If the salt, thus made, be exposed to a strong heat, all the ammonia is expelled, and phosphate of magnesia remains.

*Carbonate of magnesia* is prepared, by decomposing sulphate or muriate of magnesia by carbonate of potash or soda. The latter does better than the former, as it is usually obtained in a purer state. One hundred and forty-four grains of crystallised carbonate of soda, decompose one hundred and twenty of the crystallised sulphate of magnesia. The salts must be, separately, dissolved in six times their weight of boiling water. In this process, the carbonic acid of the carbonate unites with the magnesia of the sulphate, forming carbonate of magnesia, which is precipitated, while the sulphuric acid combines with the soda, and remains in solution. The mixture should be boiled for a short time before it is allowed to cool, and the bulky precipitate of the carbonate washed with pure water on a linen filter, in order to remove any adhering sulphate. This compound consists of carbonic acid 22 parts, and magnesia 20 parts.

Carbonate of magnesia frequently contains some sulphate of lime, which Dr Percival attributes to the use of *hard water*, in its preparation. Occasionally there are traces of silicious earth, derived from the alkali used in the ordinary process. We detect the sulphate of lime, by boiling a sample of the powder in water, and assaying it with a barytic and oxalic test; the first to show the presence of sulphuric acid, and the latter to indicate the lime.

The *incompatibles* of this carbonate are, *acids* and *acidulous salts*, *alkalis* and *neutral salts*, *alum*, *cream of tartar*, *nitrate*



of silver, acetate of mercury, corrosive sublimate, sugar of lead, sulphates of zinc, copper, and iron.

Magnesia is much used in the treatment of calculous affections, especially to prevent the formation of *uric* or *red sand*, and to carry it off when it already exists. It is productive of mischievous effects, in cases of *white sand*, because it only tends to increase the proportion of one of the ingredients of that species of gravel. The *red sand* consists chiefly of uric acid, the production and increase of which seem to be prevented by the administration of magnesia. The *white sand* is an *ammoniaco-magnesian phosphate*, and it is easy to see how the free use of magnesia might increase the mischief already existing.

Carbonate of magnesia requires nearly 2500 parts of cold, and 9000 of hot water, for solution. In many cases, it will be advantageous to give it in milk, in which it is rather more soluble.

The famous patent medicine, called *Dalby's carminative*, owes its virtues, chiefly, to the carbonate of magnesia. A few grains of the carbonate, in mixture with a little mint water, will generally answer quite as good a purpose.

*Bi-carbonate of magnesia* is a compound of two equivalents of carbonic acid and one of magnesia. It is prepared, by passing a stream of carbonic acid gas through carbonate of magnesia suspended in water; it is soluble in fifty parts of water, and the solution has been sold under the names of *ærated magnesia water* and *liquid magnesia*.

To prepare a small quantity of this liquid, put some carbonate of magnesia, finely powdered, into a bottle, and fill it nearly two-thirds full of carbonic acid water. Exp. Shake the mixture well and then filter; the liquid that passes the filter, will be found to hold magnesia in solution.

If a part of this clear solution be heated, the excess of carbonic acid will escape; the liquid will become quite turbid, owing to the precipitation of the insoluble carbonate.

## ORDER THIRD.

## OF GEOFIABLE METALS,

OR THOSE WHICH CONSTITUTE THE BASES OF THE EARTHS PROPER.

## CHAPTER I.

## Of Aluminium.

THIS metallic substance was discovered by Sir H. Davy. He obtained it, by the aid of galvanism, united with iron; on throwing water on the mixture, alumine was formed. He also procured it by decomposing alumine, by means of potassium.

Wohler, availing himself of a hint given by Oersted, procured aluminium by means of the *chloride* of that metal. This compound is made by the following process. Alumina is first precipitated from alum, by carbonate of potash added in excess. This precipitate is well washed, dried, and then mixed with powdered charcoal, oil, and sugar, into a thick paste, which is to be heated in a covered crucible till all the vegetable matter is destroyed. The black matter that remains is, while hot, to be put into a porcelain tube, properly arranged in a convenient furnace; and through this, chlorine gas, perfectly dried by means of fused chloride of calcium, is to be passed into a small, tubulated, glass balloon, attached by a glass tube to the other end of the porcelain tube. When the apparatus is filled with chlorine, the porcelain tube is to be raised to a red heat, and then the chloride will begin to appear in the balloon, in a pulverulent form. A small portion, also, escapes with the oxyde of carbon that is evolved, and causes it to fume in the air. At the end of an hour, or little more, the tube will be obstructed by the chloride, and the operation may then cease. On taking the apparatus to pieces, after cooling, the chloride is found at the end of the tube, as it issues from the furnace, partly in a crystalline form, and partly in a solid, hard mass, which is easily separated. It is of a pale, yellowish-green color, semi-transparent, and in scales. Exposed to the air, it fumes a little, giving out a smell of muriatic acid, and is changed into a clear liquid. When added to a small quantity of water, it is dissolved, and considerable heat is evolved. At a heat a little above

212°, it is volatile, and fuses nearly at the same temperature. It may be kept unaltered, in pure naphtha. When heated in a glass tube, with sulphuretted hydrogen, it absorbs 30 or 40 times its bulk of that gas, and yields it again, when added to water.

If a small fragment of the chloride of aluminium be heated, in a glass tube, with some potassium a heat will be excited, sufficiently intense to break the tube. The process Exp. may be safely conducted, in a covered platina crucible, taking care to use no more than from six to ten small balls of potassium, each of the size of a pea, and not more of the chloride than will be volatilised by the reduction. A fused mass remains, of a grayish-black color. To detach this, the crucible must be plunged into a vessel of water; a slight escape of hydrogen takes place, of a very offensive odor, and a gray powder is left, which, on minute examination, in the direct rays of the sun, is evidently composed of very small metallic plates, resembling platina. This powder, sufficiently washed with cold water, and dried, is *aluminium*.

The color of this metal is tin-white, especially when compressed in a mortar, and rubbed with a burnisher. Its aspect is decidedly metallic. It is infusible at the heat for melting cast iron. It does not conduct electricity, but this may be owing to its pulverulent form.

When heated to redness in the air, it takes fire, and burns with great brilliancy into white and pretty hard alumine. When cast on the flame of a candle, it emits sparks as brilliant as those of iron burning in oxygen gas. It burns in oxygen gas with a splendor almost insupportable, and with so much heat, that the resulting alumine is fused into intensely hard masses, capable of scratching glass. It will not burn, however, even in oxygen, unless first made red hot.

Water, at any temperature, has little or no action on aluminium. At ordinary temperatures, it is not affected by sulphuric and nitric acids.

If ignited in a stream of chlorine, it takes fire, and is changed into chloride of aluminium, which is volatilised by heat. It unites with sulphur, phosphorus, selenium, arsenic, and tellurium, generally with incandescence.

There is but one oxyde of aluminium known, and that is alumine. It consists of oxygen 8 parts, or one equivalent, and aluminium 10, or one equivalent, making the equivalent of the oxyde 18.

Alumine is the plastic principle of all clays, loams, and boles. It was formerly called argil, or argillaceous earth; but as it is now procured by the decomposition of alum, it is styled alumine

or alumina. Like all the other earths, it was formerly called a simple substance.

The oriental gems, as the sapphire and ruby, contain the purest native alumine. They consist of nothing but this earth, and a small portion of coloring matter.

To obtain pure alumine, we dissolve alum in twenty times its weight of water, and add ammonia till it ceases to give a precipitate. The alumine falls, in the form of a white, bulky mass, which must be washed with water, and afterwards dried. In this state, it still contains a small quantity of potash and sulphuric acid in combination, but for ordinary purposes, it is not necessary to separate them.

To understand the nature of the process just mentioned, it must be recollected that common alum is composed of sulphate of potash and sulphate of alumine; the ammonia decomposes the latter sulphate, and unites to the acid, while the alumine is precipitated.

Alumine may also be obtained from the solution of alum, by adding the bi-carbonate of potash.

We ascertain the purity of alumine, by boiling some of it in an ounce or two of a strong solution of pure, caustic potash, diluted with an equal bulk of water. If the alumine be pure, it will be completely dissolved; and this feature serves to distinguish it from the other earths, which are not dissolved by the caustic alkalis. Pure alumine is white, pulverulent, soft to the touch, adheres to the tongue, forms a smooth paste without grittiness, is insipid, inodorous, produces no change on vegetable colors, is insoluble in water, but mixes with it readily in all proportions, and retains a certain portion with a force which cannot be subdued, except by a white heat. It is infusible in the strongest furnace heat, experiencing only a slight condensation and consequent hardness. In small quantities, it melts under the oxy-hydrogen blast. Its specific gravity is 2, in the state of powder, but is increased by ignition.

Although alumine does not affect the vegetable colors, it appears to have the properties of an acid and of an alkali. It unites with alkaline bases, as potash, lime, and barytes. It also combines with alkalis, and gives rise to salts.

Alumine that has lost its plastic character by ignition, recovers it by solution in an acid or alkaline menstruum, and being then precipitated. Alumine is widely diffused in nature. It is a constituent of almost every soil and rock. It is the basis of porcelain, pottery, bricks, and crucibles. On account of its affinity for vegetable coloring matter, it is used in the preparation of lakes, and in the arts of dyeing and calico-printing. Fullers'

earth, ochres, pipe-clays, &c., contain alumine in large quantities.

We recognise alumine by the following characteristics. It is separated from acids, as a hydrate, by all the alkaline carbonates, and by pure ammonia. It is precipitated by pure potash and soda, but the precipitate is entirely redissolved by an excess of alkali.

The only salts of alumine, worthy of notice, are the well known article, alum, and the acetate of alumine.

There are several varieties of alum, as the sulphate of alumine and potash, sulphate of alumine and ammonia, and sulphate of alumine and soda. Before we examine the varieties of alum, it is proper to notice the compound of sulphuric acid and alumina, called *sulphate of alumina*. This may be made by digesting pure clay in sulphuric acid. The salt readily crystallises in thin, soft plates, having a pearly lustre. It has an astringent taste, and is so soluble in water, as to crystallise with difficulty. When moderately heated, the water escapes, and if the heat be increased, the acid is also driven off.

*Sulphate of Alumine and Potash.* This is the common alum of the shops, of which the *rock* and *Roman* alum, of commerce, are only varieties.

The composition of alum has been a matter of dispute; not in relation to its ingredients, but as respects their relative proportions. Reid says it consists of 88 or one equivalent of sulphate of potash, 174 or three equivalents of sulphate of alumine, and 225 or twenty-five equivalents of water, making its entire equivalent 487. Brande gives its composition thus; 132 or two equivalents of sulphate of alumine, 128 or one equivalent of bi-sulphate of potash, and 225 or twenty-five equivalents of water; total 485. Phillips regards it as a compound of two equivalents of sulphate of alumine, one equivalent of bi-sulphate of potash, and twenty-two equivalents of water; making in all four hundred and fifty-eight.

We infer, from the different results of analysis in the hands of such able chemists, that true alum may be formed out of very diverse proportions of the same ingredients. This inference derives some support from the well known fact, that an alum may be formed with either potash, ammonia, or soda, joined to sulphate of alumine.

Alum is made, ordinarily,\* by adding aluminous earth to the

\* Alum is also made, in large quantities, from what are called *alum ores*. These ores contain sulphur, potash, and alumine; by a strong heat and exposure to the air, the sulphur is converted into sulphuric acid, and combines with the other ingredients, to form the salt.



residuum of the preparation of nitric acid, viz. bi-sulphate of potash; and it is obtained, frequently, in very large crystals, which are slightly efflorescent. The taste of alum is sweetish, harsh, and acidulous. A fluid ounce of cold water dissolves 30 grains; the same quantity of boiling water dissolves 4 drachms, and the solution reddens litmus paper. Alcohol does not dissolve any portion of alum. The *incompatibles* of this salt are *alkalis* and *alkaline salts*; these first neutralise the excess of acid and then throw down the alumine. It is also decomposed by *carbonate* and *muriate of ammonia*, *carbonate of magnesia*, and *tartrate of potash*, *lime water*, *sugar of lead*, *salts of mercury*, and by many vegetable and animal substances, as *kino*, *galls*, and *Peruvian bark*. Hence, powders made of alum, and either of these vegetable matters, with a view to increase the powers of the alum, fail to accomplish the design. One of the greatest mistakes, however, in the use of alum, is to combine it with the acetate of lead, to increase astringency; a double decomposition ensues, when these articles are dissolved, which destroys the medicinal character of both. The whole of the sulphuric acid of the alum goes to the lead of the sugar of lead, and an insoluble (of course inert) sulphate of lead is produced; on the other hand, all the acetic acid of the sugar of lead joins the alumine and potash of the alum, and a compound results which is void of astringency.

Alum possesses the property of retarding, and in some cases, of preventing, the acetous fermentation. Paper-hangers practice on this principle, and hence they always add a portion of alum to their paste, in order to obviate its tendency to become sour in hot weather. Solutions of glue may be preserved, in the same way, for a considerable length of time.

The property of clearing turbid water has long been ascribed to alum, but in no instance has its mode of acting been explained, so far as my observation extends. I have supposed, that the basis of the turbid matter is alumine, which is neutralised by the excess of acid in the alum, and that this is the secret of the change; but there may be a better explanation.

The power of alum to produce coagulation has been usefully applied; hence, we have the *alum whey*, made by boiling two drachms of water in a pint of new milk, and straining, when cold. Also, the *alum curd*, formed by agitating a drachm of alum with the white of an egg.

Besides the medicinal uses of alum, it is also employed extensively in the arts. It is added to tallow, to give it hardness. The printers of calico employ it, to considerable extent, for various purposes. Wood soaked in a solution of alum, does not

easily take fire, and paper treated in the same way, is also made, to a certain extent, incombustible, while at the same time, it excludes moisture. Crayons generally consist of the earth of alum, finely powdered, and colored.

The usual form of the crystals of alum is octahedral, but if an excess of potash be present, they assume the form of cubes, and hence the name *cubic alum*.

*Alumen exsiccatum*, dried or burnt alum, is made by heating alum on an iron shovel, over red hot coals. It gradually swells, and undergoes the watery fusion; it loses more than a third of its weight, which loss is occasioned by the escape of its water of crystallisation. If the heat be very intense and long continued, a great part of the acid of the salt is driven off. This preparation has been highly applauded, in the treatment of colica pictonum.

When alum is ignited with charcoal, a spontaneously inflammable compound results, which has long been known under the name of *Homberg's pyrophorus*. The potash is probably decomposed in this process, as well as the acid of the alum, and pyrophorus is, by some, supposed to be a compound of sulphur, charcoal, potassium, and alumine.

A pyrophorus is easily prepared, by mixing alum with sugar, and exposing it to heat. Different proportions have been recommended; three of alum to one of sugar do very well. Two or three ounces of the mixture should be melted over an open fire, in an iron ladle, continuing the heat till all the water of crystallisation of the alum has been completely expelled, and the mass has assumed such a consistence, that it may be easily reduced to powder; it must be stirred, constantly, with an iron rod or spatula, and prevented from adhering to the sides of the iron ladle. Let it then be reduced to Exp. powder and exposed to a red heat, in a small iron bottle or tube retort, till the gas which is disengaged begins to burn with a pale blue flame, a short time after which it is to be removed from the fire, and the mouth of the vessel plugged up with a little well worked clay, till it becomes cold, replacing it then with a cork, and making it air-tight with gas lute.

This substance inflames in atmospheric air at natural temperatures, and burns brilliantly in oxygen. Small quantities should be used at a time, and the cork of the iron bottle replaced, whenever as much as may be required is taken out, as it is soon destroyed by exposure to the air. Part of the carbonaceous matter of the sugar decomposes the sulphuric acid, the potash, and the alumine during its preparation, so that the product may be regarded as a sulphuret of aluminium and potas-

sium, mixed with the rest of the carbon of the sugar; many are inclined to believe, that a small quantity of potassium is set free, and that the inflammable properties depend more upon this, than upon the minute state of division in which the carbon is obtained.

We have said, that soda has been used in the formation of alum, in the room of potash, and hence the term, *soda alum*. This variety was first made by Dr Ure, and is described by him, in the Quarterly Journal of Science, (England,) for July, 1822. It consists of 3 equivalents of sulphate of alumine, or 174, and one equivalent of sulphate of soda, or 72, with 25 equivalents of water, or 225.

The form and taste of soda alum, are the same with those of common alum; but it is not so hard, being easily crushed between the fingers, to which it gives an appearance of moisture. Its specific gravity is 1.6. One hundred parts of water at 60°, dissolve 110 parts of this species of alum, while the same quantity of water dissolves only 9 parts of ordinary alum.

The only injurious contamination of alum, results from an admixture of sulphate of iron. This is detected by the ferrocyanate of potash. To get rid of it easily, M. Thenard advises to dissolve the alum in boiling water, and to agitate the solution with rods, as it cools. In this way, the salt is reduced to a fine, granular powder, which, by several washings with cold water, yields a perfectly pure alum.

*Acetate of alumine* is an important article in the hands of the dyer. It may be made, by digesting freshly made alumine in strong acetic acid, but this would be too expensive for ordinary use. The calico printers prepare it, by the double decomposition of alum and sugar of lead, whence result the sulphate of lead, an insoluble powder, and acetate of alumine which remains in solution. But it is a more economical mode, to use the acetate of lime, instead of the acetate of lead. The acetate of lime can be had at little cost, by adding lime to pyroligneous acid, which contains a large proportion of acetic acid. When the acetate of lime and sulphate of alumine and potash meet, in solution, a soluble acetate of alumine is formed, and a precipitate of sulphate of lime is thrown down, which is nearly insoluble. By careful evaporation, the solution yields fine crystals, which quickly deliquesce.

The acetate of alumine is composed of 51 acetic acid, or one equivalent, and 18, or one equivalent of alumine. The affinity between its component parts is very feeble; hence, the cotton fibre, moderately heated, has the power of decomposing it.

## CHAPTER II.

## Of Silicium.

BERZELIUS was the first person who obtained this metallic base. He fused silica with charcoal and iron, in a blast furnace, and procured an alloy of iron and *silicium*, which, by the action of a diluted acid, gave more hydrogen than the same weight of iron would have yielded. If seven parts of iron, five of silica, and from one-fourth to eight-tenths of a part of soot be treated as above, the same result will be obtained. From the results of the action of diluted acids on the alloy, Berzelius infers, that silica consists of about 46 parts silicium and 54 oxygen, in 100.

Berzelius obtained silicium in an insulated state, by attempting to reduce fluoric acid, by means of potassium. If pure potassium be burned in silicated fluoric acid gas, the latter is condensed, and a brown substance is separated. This, after being boiled in water and dried, burns in oxygen gas, evolving no carbonic acid, but yielding silicated fluoric acid and silica.

If potassium be heated in a glass tube, together with dry, silicated fluuate of potash, the silica is decomposed, its oxygen going to the potassium, and silicium being produced, in union with some hydrogen, which may be detached by heating the mixture in a crucible.

Pure silicium is incombustible, even in oxygen gas. It is not acted on by water, aqua regia, sulphuric and nitric acids, nor by caustic alkalis; but fluoric acid, especially if a little nitric acid be added, has a solvent action on it. When it is heated with nitrate of potash and a small quantity of carbonate of soda, detonation instantly ensues. When the vapor of sulphur is brought in contact with ignited silicium, the latter suddenly becomes incandescent. The resulting sulphuret decomposes water with extreme rapidity; sulphuretted hydrogen is evolved, and the water dissolves so much of the newly regenerated silica, that it thickens and coagulates during evaporation, and deposits a gummy, transparent mass.

Silicium takes fire in chlorine gas, at a red heat, and a yellowish liquid is formed, having an odor like that of cyanogen. This odor is highly volatile, and when water is added, gelatinous silica is deposited.

If silicium be heated with carbonate of potash, a brilliant inflammation is seen, and carbonic acid is evolved. With hy-

drates of the fixed alkalis, it detonates and produces a vivid incandescence, at a temperature below redness.

Silicium is a non-conductor of electricity. It does not combine easily with other metals; but if silica be reduced in connexion with other metals, it readily yields alloys of silicium. From the most careful examination, it is agreed that silica is composed of 8 silicium, or one equivalent, with one equivalent or 8 of oxygen.

*Silica* or the oxyde of silicium, is a very abundant natural product. It constitutes a principal ingredient of several rocks, that form extensive mountain masses, such as granite. It composes a large proportion of the sand found in the beds of rivers, and on the seashore. It enters largely into several minerals of rarer occurrence, as rock-crystal, quartz, flint, chalcedony, cornelian, opal, &c. It also exists in fertile soils, rendering them porous and open to the transmission of moisture. In the arts, its most important use is to make glass, and in small quantities, to improve porcelain.

Mr Brande says, that silica, pure enough for most purposes, may be made, by heating rock crystal to redness, quenching it in water, and reducing it to powder. It is also made, by calcining common gun-flints, at a red heat. When reduced to powder, they are mixed with four times their weight of carbonate of potash, and the mixture fused in a crucible, exposed to a red heat. The crucible must be much larger than it need be simply to hold the materials, because these effervesce, and sometimes run over. After the whole has been melted, the mass is to be poured on a copper or iron dish. The compound thus obtained, consists of alkali and silicious earth. Dissolve it in water, filter the solution, and pour it into dilute sulphuric or muriatic acid. The precipitate that falls must be collected, and washed with hot water, until that fluid comes off tasteless, after which it must be dried.

Pure silica is a light, white powder, insipid, inodorous, and giving to the fingers a sensation of roughness. It is fixed in the fire, and infusible by ordinary heat. It fuses, however, before the oxy-hydrogen blast, with greater facility than lime or magnesia.

In its solid form, silica is quite insoluble in water, and hence does not change vegetable blue colors. Berzelius says, that when silica, in its nascent state, comes in contact with water, it is dissolved freely, and on evaporating the fluid, a hydrate of silica is obtained. This hydrate is partially decomposed by a very moderate heat, but a red heat is requisite to expel the whole of the water.

Silica appears to possess the characters of an acid, rather than



of an alkali. No acid acts on it, excepting the fluoric or hydrofluoric; but it is dissolved by solutions of the fixed alkalis, and combines with many of the metallic oxydes. Hence, *silica* is sometimes called *silicic acid*, and its compounds, *silicates*.

Fuse three parts of white sand (purified previously by washing it with water and exposing it to a red heat) with one part of purified carbonate of potash, in a crucible, placing it in a furnace so that it may be exposed to a very strong heat. A transparent, brittle glass is obtained, quite insoluble in water, and having usually a greenish tinge, derived from a minute portion of iron in the materials or in the crucible. The sand and the carbonate should be well mixed together, and the crucible should not be filled more than a third full, placing it on a piece of brick, so as to rest an inch and a half above the branders of the furnace; a cover should also be placed loosely over it. The glass is formed, by the silica of the sand uniting with the potash, the carbonic acid being disengaged. About 800 or 1600 grains of materials give a very good specimen of glass. Exp.

Fuse another portion of the same materials with one or two grains of the peroxyde of manganese. The glass will be found to have acquired a light tinge of purple, and to be free from the green or greenish-yellow tinge which the iron communicates. Exp.

In another crucible, fuse three parts of the carbonate of potash with six of the red oxyde of lead and ten of sand. The mixture will melt at a much lower temperature than the former, and the glass formed, is of the same nature as the flint glass prepared at the different glass works. The same materials, with the addition of a small quantity of manganese, give a rich purple colored glass, and with the oxyde of cobalt it may be obtained of a fine blue color. Exp:

Fuse one part of sand with three of the carbonate of potash, and pour the liquid on an iron plate, after effervescence has ceased. The compound, procured in this manner, has the same appearance as glass, but it deliquesces speedily on exposure to the air, and is very soluble in water;\* it is usually called *silicated potash*. In this case, all the carbonic acid of the alkali is not expelled, and its solution effervesces on the addition of acids. It must be kept in close vessels. Exp.

If some sulphuric or muriatic acid be added to a solution of silicated potash, the latter will be decomposed. The acid unites with the alkali, and the silica is precipitated, in combination with part of the water. Exp.

\* The watery solution is called *liquor silicum*, or liquor of flints.

If we add a solution of the muriate of ammonia to a solution of the silicated potash, the muriatic acid of the former  
 Exp. will unite with the potash, and all the silica will be precipitated.

Crown or window glass is made of kelp and fine white sand. Bottle glass is made of impure materials, as the refuse of the salt employed by soap makers and common sand.

Flint glass is quite insoluble in water in its ordinary state, but it may be dissolved in small quantity, if reduced to a very fine powder; and the solution turns the infusion of blue cabbage to a green.

## CHAPTER III.

Of Glucinium, Yttrium, Zirconium, and Thorium.

THESE are the metallic bases of the earths *glucina*, *yttria*, *zirconia*, and *thorina*. Glucina is found in the beryl and emerald; yttria was discovered by Gadolin, in the mineral called Gadolinite; zirconia was found by Klaproth, in the zircon of Ceylon; and thorina was obtained a very short time ago (1829) by Berzelius.

As none of these earths are of any practical importance, I shall not occupy more space in their consideration.

## ORDER FOURTH.

OF METALS WHOSE OXYDES CANNOT BE  
 REDUCED BY HEAT ALONE.

## CHAPTER I.

Of Iron.

THE equivalent or combining number of iron is 28, and its specific gravity 7.7. It has a peculiar gray color, and strong metallic lustre, which may be increased by polishing. It is inferior in ductility and malleability to several metals, but ex-

ceeds them all in tenacity. Heated to redness, it becomes soft and pliable, so that it may be beaten into any form, and incorporated or *welded* with another piece of red hot iron, by hammering. It is this property, that makes iron so important in the arts, and for many practical purposes in common life.

Iron has a strong attraction for oxygen. If the atmosphere be perfectly dry, it undergoes very little change, but when moisture is present, it oxydises or *rusts*, in the course of a few days. When heated to redness in the open air, it absorbs oxygen rapidly, and is coated with black scales, called the *black oxyde* of iron. In an atmosphere of oxygen gas it burns rapidly, if introduced in the state of fine wire. It decomposes the vapor of water and unites with its oxygen, while the hydrogen is liberated.

The filings and turnings of iron are used in the practice of medicine. They should be purified by the application of the magnet, as they are often mixed with copper and other metals. Iron wire should be preferred for making the filings, because none but the purest metal can be drawn into wire, without great difficulty.

*Oxydes of Iron.* Several combinations of oxygen and iron have been noticed by different writers; but we design to specify but two, as all others appear to be modifications of them.

The *protoxyde of iron* is composed of one equivalent of each of its components, viz. oxygen 8, and iron 28. It is not easily procured in a pure state, from the great affinity which it has for oxygen, and the facility with which it attracts an additional portion of this element, even from atmospheric air, when moistened. This compound, however, is the principal product of the combustion of iron in oxygen gas, and it may be procured, also, by transmitting watery vapor over metallic iron, in the manner directed for the preparation of hydrogen gas. The *black oxyde of iron*, that falls in scales from iron when it is at a high temperature and hammered on the anvil, appears, from recent experiments, to be composed of the protoxyde and peroxyde of this metal.

Add a solution of potash or soda to a solution of the sulphate of iron; if it contains no peroxyde, a precipitate of a white color is thrown down, which is a compound of the protoxyde of iron and water. It soon becomes green, however, and ultimately assumes a red color on exposure to the air, attracting an additional quantity of oxygen.

To prepare *peroxyde of iron*, a crucible may be filled with green sulphate of iron, and exposed to a good red heat, in a furnace or open fire, for one or two hours, luting a cover on with a little clay previously, and leaving a small aperture for the

escape of gaseous matter. The water of crystallisation in the green sulphate, is completely expelled with part of the sulphuric acid, the rest of which is decomposed, communicating part of its oxygen to the protoxyde of iron, while sulphurous acid is disengaged. It is of a red color, very dark at first, but becoming lighter when it cools.

It is usual to say, that the equivalent of the peroxyde is 40, and that it is composed of 28 iron (one equivalent) and 12 oxygen (or one equivalent and a half). But it seems more correct to say, that its equivalent is 80; then its composition would be 56, or two equivalents of iron, and 24, or three equivalents of oxygen.

The protoxyde is the base of the native carbonate of iron, and of the green vitriol of commerce. It has a dark blue color, and when melted with vitreous substances, gives them a blue tint. The ferrocyanate of potash added to it gives a white precipitate, and solution of gall-nuts produces no change.

The peroxyde exists as a natural product, under the name of *red hematite*. Fused with vitreous substances, it communicates to them a red or yellow color. It combines with most acids, forming salts, the greater number of which are red. With ferrocyanate of potash, it forms Prussian blue, and herein is an important distinction between the two oxydes, especially in a manufacturing point of view. Infusion of gall-nuts, added to the peroxyde, gives a black color, and the sulphocyanate of potash occasions a deep blood-red. Hence, in testing a solution for iron, if gall-nuts and ferrocyanate of potash produce no change, it may be inferred, that the iron is in state of protoxyde. If a few drops of nitric acid be added, the protoxyde will be changed to peroxyde, and a deep blue will be instantly formed.

That the *black oxyde* is really a compound of the oxydes already noticed, is rendered probable by the action of acids. If the black oxyde be digested in sulphuric acid, an olive colored solution will be formed, containing two salts, viz. sulphate of the peroxyde, and sulphate of the protoxyde, which may be separated from each other by means of alcohol, which dissolves the per salt, but has no action on the other. The black oxyde of iron is the cause of the dull green color of bottle glass.

A black oxyde was formerly in use, called *martial æthiops*, made by exposing iron filings, moistened with water, to the open air, at a temperature of 70°. The *colcothar* of the old books, is the same with the *peroxyde* of iron, noticed above. It was formerly employed as a tonic. The same oxyde was known to

the ancient chemists, by the term *crocus martis astringens*; whereas, an oxyde produced by moisture, was called *crocus martis aperiens*, and, also, the *saffron of Mars, by dew*.\*

*Sulphurets of Iron.* Two varieties of the combination of sulphur and iron are found in nature, viz. the *protosulphuret* and the *bi-sulphuret*.

*Protosulphuret of iron* is, as its name imports, a compound of one equivalent of sulphur and one equivalent of iron, viz. 16 and 28, making the equivalent of the compound, 44. It constitutes the magnetic iron pyrites, found in many sections of our country. It is a brittle, yellow substance, of metallic lustre and feebly attracted by the magnet. By exposure to the air and moisture, it is gradually changed into the protosulphate of iron.

*Bi-sulphuret of iron* is composed of two equivalents of sulphur, 32, and one of iron, 28. It is the common iron pyrites, so abundantly diffused on the face of nature, in many parts of the world, and from which large quantities of the green copperas of commerce, are manufactured. When heated to redness, it loses half its sulphur, and is converted into the protosulphuret. It is insoluble in sulphuric and muriatic acids.

*Iodide of iron* may be made, by heating iron in the vapor of iodine. It is converted into hydriodate of iron, by the action of water, as the following diagram shows.

152 or 1 equiv. of	{ iodine	124	—	125 hydriodic acid.
iodide of iron	{ iron	28		
9 or 1 equiv. of	{ oxygen	8	—	36 oxyde of iron.
water	{ hydrogen	1		
161		161		161 hydriod. of iron.

*Phosphuret of iron* may be formed, by heating the phosphate of iron with charcoal. It is sometimes contained in metallic iron, to which it proves injurious, by making it brittle at ordinary temperatures.

*Carburet of Iron.* There are several compounds of carbon and iron, that are of great importance in the arts. Thus, cast or pig iron, steel, cast steel, and the common plumbago, are so many species of carburet of iron.

It is well known, that the native oxydes of iron are blended with argillaceous and silicious matters, and that their reduction to the metallic state by the agency of charcoal and lime, at a high temperature, is a matter of every-day occurrence. The

\* See Macquer, vol. i. page 367.



oxyde of iron is, for the most part, decomposed in this operation; its oxygen unites with one portion of the carbon employed in the reduction, while the iron unites with another, the products being, in the one case carbonic acid, and in the other, carburet of iron. The earthy matters unite with a little undecomposed oxyde of iron, and give rise to a vitreous mass, called *slag*.

The carburet thus formed, is void of ductility and malleability, and is very brittle. At a red heat, it enters rapidly into fusion, and cannot be welded. It contains a forty-third part of its weight of carbon, some manganese, and the metallic bases of several earths, with other occasional admixtures.

The combined action of air and intense heat changes cast iron into malleable iron, if a reverberatory furnace be employed. In the operation, the whole of the undecomposed ore is reduced, earthy matters rise to the surface, and the carbon is oxydised. In proportion to the purity of the iron, its fusibility diminishes, until, at length, the iron becomes solid. Then, while it continues hot, it is subjected to the operation of the roller or hammer, by which its particles are more condensed. By these processes, the metal loses nearly all its carbon, and becomes almost pure. Berzelius has shown, that even in this state, it contains one half per cent. of carbon, and a little silicium.

*Steel* is common iron, altered by combination with carbon. Bars of the purest malleable iron, are surrounded with powdered charcoal, and exposed to a red heat, for a considerable length of time. In this process, the iron unites with a hundred and fiftieth part of its weight of carbon, and acquires new properties. The carburet, thus formed, is less ductile and malleable than iron, while it exceeds it in hardness, sonorousness, and elasticity. Its texture is more dense than that of iron, and it admits of a finer polish. It does not fuse at a strong, red heat, and it may be welded with iron. When combined with a further dose of carbon, it forms *cast steel*. This is harder, more elastic, denser, and susceptible of a higher polish than common steel. But, on account of its greater fusibility, it cannot be welded.

The difference between *cast iron* and *steel*, is really chemical; the latter contains purer iron and less carbon than the former. We readily distinguish steel from malleable iron, by the action of an acid. If a drop of nitric or muriatic acid be put on steel, and the surface be washed with water, in a few minutes afterwards, a black spot will be seen on the surface. This results from the acid acting on the iron, while the carbon remains unchanged. The same appearance is presented with cast iron, which contains still more carbon than steel.

From these circumstances, it is obvious, that when steel or cast iron is used for making hydrogen gas, a smaller quantity will be obtained, from a given weight of materials, than when forged iron is used, and that the product will be contaminated, either with carbonic acid or carburetted hydrogen gas.

The common *plumbago* or *black lead* of the shops, (known also as graphite,) is a native carburet of iron, containing ninety-five per cent. of carbon. The air exerts no action on it, and it powerfully resists almost all chemical substances. This property may be attributed to the charcoal contained in it, a substance which is known to withstand the action of more agents than any other substance. The name, *black lead*, was given to this carburet, on account of its iron-gray color, and its use in the form of pencil. It has a decided metallic lustre, and granular texture; is largely employed in the manufacture of crucibles, and also to burnish iron, in order to secure it from rust.

*Muriate of iron* may be formed, by putting iron filings or turnings into muriatic acid, diluted with half its bulk of water. Part of the water is decomposed; the oxygen goes to the iron, forming a protoxyde of iron, which is dissolved by the acid, and hydrogen is disengaged. The solution, thus made, has a green color; and when exposed to the air, it attracts oxygen, the iron is changed to a permuriate, part of the oxyde being at the same time deposited.

If some peroxyde or rust of iron, be digested in about four times its weight of muriatic acid, a permuriate is obtained in solution. It is this which is used in preparing the *muriated tincture of iron*. The protomuriate is insoluble in alcohol, while the permuriate is quite easy of solution in that menstruum.

The protomuriate of iron is composed of 37 muriatic acid and 36 protoxyde of iron; in the permuriate, the acid is the same; while the equivalent of the peroxyde is 40.

The *tinctura ferri muriatis* or muriated tincture of iron, is largely used in the practice of medicine. Its color is brownish-yellow; it has a styptic taste and a very peculiar odor, of rather an ethereal quality.

The *incompatibles* are, *alkalis* and their carbonates, *infusions of astringent vegetables*, and *mucilage of gum arabic*.

If a solution of the muriate of iron be evaporated to dryness, the oxygen of the oxyde of iron will unite with the hydrogen of the muriatic acid, while the chlorine of the latter, joining the metal, gives rise to a *chloride of iron*.

When muriate of iron, in solution, is mixed with muriate of ammonia, and evaporated to dryness, a compound is obtained, which has been called the *muriate of ammonia and iron*. It

has also been known by the names of *ammoniated iron*, *flores martiales*, *ens martis*, &c.

Some direct it to be made, by subliming a mixture of muriate of ammonia and muriate of iron, or the muriate of ammonia and peroxyde of iron. Paris says, 'the composition of this article is very variable, depending upon the degree of heat, and length of time employed for its preparation.' It is no longer employed in medicine.

*Nitrate of iron* is composed of 54 nitric acid, (one equivalent,) and 36 protoxyde of iron, (one equivalent,) and when crystallised, it contains seven additional equivalents of water.

We may form this compound, by adding one or two ounces of iron turnings to one ounce of nitric acid, in a glass flask or bottle, previously diluting the acid with seven or eight ounces of water. The iron is slowly oxydated, and then dissolved; and the solution being evaporated with a gentle heat, yields green crystals of nitrate of iron. They are of very little use excepting, occasionally, in the laboratory.

If the acid is not diluted with so much water, the iron attracts a larger quantity of oxygen, and the liquid acquires a reddish-brown color, *pernitrate of iron* being formed, which does not crystallise, on evaporating the liquid.

When only a small quantity of water is added to the acid, and the iron in a minute state of division, the action is very turbulent. Put three or four drachms of nitric acid into a glass, add two drachms of water, and throw into it 200 or 300 grains of iron filings. Part of the acid is immediately decomposed, the iron is oxydated and combines with another portion of acid, and a large quantity of fumes is disengaged, which must be carefully avoided, as they are composed principally of nitrous acid and nitric oxyde; what remains, consists of nitrate of iron, with an excess of acid.

*Sulphate of iron* is the article that next claims our attention. It is known by the various names of *green vitriol*, *sal martis*, *green copperas*, *vitriolated iron*, &c.; but the appellation which best accords with the true spirit of our nomenclature is, the *protosulphate of iron*, or the *sulphate of the protoxyde of iron*. It is composed of 40 sulphuric acid, (one equivalent,) and 36 protoxyde of iron, (one equivalent); the crystallised sulphate contains seven additional equivalents of water, or 63.

This salt is largely prepared for commercial purposes, by exposing the native sulphuret of iron to air and moisture, the iron being converted into an oxyde and the sulphur into sulphuric acid, by attracting oxygen. On the small scale, it may be prepared, by mixing six parts of iron with ten of sulphuric

acid, and sixty of water, evaporating the solution in a glass or earthen vessel, after the effervescence arising from the disengagement of hydrogen gas has ceased, till a rod dipped into it, presents appearances of crystallisation, when held in the air. The solution may then be filtered, and green crystals of the sulphate will be formed as it cools. The reaction that takes place between the iron, the sulphuric acid, and the water, has been already explained. The crystals contain nearly seven-eighths of their weight of water, and are soluble in two parts of cold, and in less than their weight of boiling water.

Pour some strong sulphuric acid on iron filings in a glass flask. Instead of a turbulent action taking place, as in the preceding instance, scarcely any traces of decomposition are observed; if, however, the flask is now exposed to heat, (a chauffer is better than a lamp in the present instance,) the iron takes oxygen from part of the acid, and sulphurous acid is disengaged, the oxyde of iron combining with another portion of the sulphuric acid, which is not decomposed, and forming sulphate of iron. This process is never adopted for preparing the sulphate, but merely for experimental illustration.

Sulphate of iron is sometimes required free from its water of crystallisation, as in the preparation of strong acetic acid. For this purpose, it may be exposed to heat in an earthen vessel, till it becomes quite dry and of a greenish-white color; almost all the water is expelled in this manner, and the salt becomes, at the same time, quite opaque; the heat must be applied by placing it over a furnace or good chauffer, taking care, however, not to allow it to be so great as to expel any of the sulphuric acid; otherwise it will acquire a red color, from the formation of some peroxyde of iron.

Dissolve some of the green crystals of the sulphate in spring water; the solution appears quite turbid from the free oxygen which it contains, converting part of the protoxyde of iron, in the sulphate, into peroxyde of iron,\* a portion of which is deposited; the acid that is sufficient to retain a given quantity of iron in the state of protoxyde in solution, not being able to dissolve it all when it acquires an additional portion of oxygen. Drop a little sulphuric acid into the solution and it will become quite clear, the precipitated oxyde being dissolved again. Exp.

Boil a solution of an ounce of the crystallised green sulphate of iron in two or three ounces of water with a drachm or two

\* Called, sometimes, oxysulphate of iron. The use of water that has been boiled, would obviate this difficulty.

of nitric acid, and evaporate the solution to dryness in an earthen vessel. Whenever the nitric acid is added to the solution, it assumes a very dark color, and the dry mass which is afterwards obtained, consists principally of the persulphate of iron, the oxyde having attracted an additional portion of oxygen from the nitric acid. On digesting it in water, the persulphate is dissolved, and a solution of a reddish-brown color is obtained.

A similar solution may be obtained, by exposing a solution of the green sulphate to the air for a long time, part of the peroxyde that is formed being deposited.

The crystals of sulphate of iron are transparent, and of a fine green color. When exposed to the air, they effloresce, and are coated with a yellow powder, which results from the absorption of oxygen. Exposed to heat, they undergo watery fusion, and at a higher temperature, the acid is driven off, and the peroxyde of iron alone remains, which is known by the name of *coelesthar*. The acid obtained in this operation, prior to the formation of the peroxyde, was called the *glacial oil of vitriol*, because it solidified at common temperatures.

Sulphate of iron is soluble in two parts of water at 60°, and three-fourths at 212°. The solution reddens vegetable blues. It is insoluble in alcohol, but if the iron in the salt gain an increase of oxygen, it becomes soluble in that menstruum.

The *incompatibles* are numerous. Thus, every salt, whose base forms an insoluble compound with sulphuric acid, is incompatible; the *earths*, *alkalis* and their *carbonates*, *borax*, *saltpetre*, *sal ammoniac*, *tartrate of potash and soda*, *acetate of ammonia*, *nitrate of silver*, *sugar of lead*, and the *soaps*, are all incompatible. Some have added, also, the *astringent vegetable solutions*, because they strike a black color, and seem to occasion new compounds; but Brande and Paris are of opinion, that there is no objection to this mode of administration.

*Carbonate of iron* is a compound of 22 carbonic acid, and 36 protoxyde of iron. It is sometimes called *subcarbonate of iron*, *precipitated carbonate*, and *rubigo ferri* or *the rust of iron*.

This article may be obtained, by dissolving 144 grains of the crystallised carbonate of soda, and pouring the solution into a vessel containing 139 grains of crystallised green sulphate of iron, dissolved in one ounce of water, and collecting the precipitate on a filter. The reaction that ensues, may be understood by the following diagram.



Before decomposition.		After decomposition.	
54 or 1 equiv. of carb. soda	{ soda 32	72 sulph. of soda.	
	{ carb. acid 22		
76 or 1 equiv. of protosul. iron	{ sulph. acid 40	58 carb. of iron.	
	{ protox. iron 36		
<hr/> 130	<hr/> 130	<hr/> 130	

The carbonate of iron (or rather the protocarbonate or carbonate of the protoxyde of iron) is thrown down, and the sulphate of soda remains in solution. The precipitated carbonate should be washed on a filter, with hot water, recently boiled, to expel the air usually contained in it. It quickly attracts oxygen from the air, and changes from a dingy green to the brown color of the rust of iron, losing also some of its carbonic acid.

If we take some of the precipitated carbonate as soon as it has been washed with hot water, and put it, while moist, into a bottle half full of carbonic acid water, and shake them well together, we shall have a bi-carbonate of iron, <sup>Exp.</sup> which possesses all the valuable properties of the carbonated chalybeate waters, and in which we may detect the iron by the usual tests. If a portion of this solution be heated in a common oil flask, the excess of carbonic acid will escape, and carbonate of iron will again be thrown down.

Dr Ure says, the best carbonated chalybeate springs may be closely imitated, by dissolving 3 grains of sulphate of iron and 61 of bi-carbonate of potash, in a quart of cool water, in a close vessel, and frequently agitating its contents.

Expose some iron filings to a red heat, in order to decompose any oily matter that may adhere to them, and then put a portion of them into a small flask or bottle, and pour in carbonic acid water, till it is about two-thirds full. On <sup>Exp.</sup> shaking it for a few minutes, it will be seen that a part of the iron is dissolved. In this case, the iron is oxydated at the expense of a portion of water, which is decomposed, and it is then dissolved by the carbonic acid. Water entirely free of carbonic acid, does not decompose iron at common temperatures. Carbonate of iron is sometimes made, by exposing iron filings to the action of air and water. The iron becomes oxydated, and the oxyde then absorbs carbonic acid.

It is usually of a chocolate-brown color, and is generally seen in small, conical lumps. It has no odor, and its taste is slightly styptic. It is insoluble in water, as we have seen, but acids act

on it, occasioning effervescence. The latter are, of course, *incompatible*, and cannot be conjoined with it, so as to avoid decomposition.

The *liquor ferri alkalini* is regarded, by Reid and some others, as depending for its virtues on the formation of bi-carbonate of iron. It has been affirmed to be the same with the *tinctura martis alkalina* of Stahl, but Brande holds a different opinion. The formula for its preparation is as follows.

Take of iron filings, two and a half drachms;

Nitric acid, two ounces;

Distilled water, six ounces;

Solution of carbonate of potash, six ounces.

The acid and water, being mixed, are poured on the iron, to form a nitrate of iron. After the action has ceased, pour off the acid liquor, and add it gradually to the solution of carbonate of potash, stirring occasionally, until a brown-red color appears, and effervescence has ceased. Then set it aside, for several hours, to settle, after which let the solution be poured off and bottled.

Its exact chemical composition has not been ascertained, and as it has ceased to be employed, it is no longer an object of importance.

*Phosphate of Iron.* I introduce this salt, because it has acquired, of late, some importance, by reason of the opinion of Carnichael and others, that iron, united to an *animal* acid, enters the system in larger quantity, and more intimately combines with the fluids, than when it is combined with a mineral or vegetable acid.

If equal parts of phosphate of soda and sulphate of iron be mixed together, in solution, a blue precipitate will be  
Exp. formed, called phosphate of iron. In this case, there is plainly a double decomposition, or a mutual exchange of base.

Phosphate of iron, in this state, is a tasteless powder, insoluble in water, but soluble in nitric acid. It occurs native, in blue colored, prismatic crystals.

*Acetate of iron* may be formed, by digesting iron in diluted acetic acid, or by adding a solution of sulphate of iron to a solution of acetate of lead. Here, also, is a double decomposition, the results of which are, an insoluble precipitate of sulphate of lead, and the acetate of iron, which is held in solution.

The tincture of the acetate is sometimes directed to be made, by mixing two parts of acetate of potash with one of sulphate of iron, in an earthen mortar, till a soft mass is obtained; then digesting in alcohol, which takes up the acetate of iron, and leaves sulphate of potash behind.

The *incompatibles* of this salt are, the *alkalis*, *alkaline earths*, and their *carbonates*, most of the *acids*, and in general, such articles as are incompatible with other acetates.

The triple salt called *tartrate of potash and iron*, is a compound whose precise chemical nature is not uniformly agreed on. Mr Phillips regards it as being, generally, a mixture of metallic iron with cream of tartar, colored by the oxyde of iron. He thinks, however, that when carefully prepared, it may be a true triple salt. Brande calls it a *pertartrate* of iron and potash, while Reid makes it to be a *prototartrate*. The latter chemist gives its composition thus; tartrate of the protoxyde of iron, 102, or one equivalent, and tartrate of potash, 114, or one equivalent, making the total equivalent 216. The best mode of preparing this salt is as follows. Mix two parts of the supertartrate of potash and one of iron filings intimately together, and expose the mixture for 15 or 20 days to the air, moistening it frequently with water. The iron is oxydated, and, combining with the excess of acid in the cream of tartar, forms tartrate of iron, which remains in combination with the neutral tartrate of potash that is left; and on boiling it for a few minutes in four times its weight of water, it is dissolved. It may be obtained in the solid form by filtering and evaporating the solution, but it does not crystallise. This compound is known in the books by the names of *tartarised iron*, *tartrate of iron*, &c. It is a brownish-green colored substance, destitute of odor, and having a slightly styptic taste. It attracts humidity from the atmosphere, and yet, according to Paris, it does not deliquesce. It is very soluble in water, and the solution may be kept for a considerable length of time, without undergoing any change, excepting that of depositing some tartrate of lime, which occurs as an occasional impurity in the cream of tartar. The *incompatibles* of this salt are, all *strong acids*, *lime water*, *hydrosulphuret of potash*, and *astringent vegetables*. The *fixed alkalis* and their *carbonates* decompose the solution of this salt, very slowly, except heat be applied; but *ammonia* and its *carbonate* have no action on it, and hence we may exhibit iron in solution with an alkali, without risking any precipitation.

It should be remarked, that there may be a tartrate of iron, distinct from this tartrate of potash and iron, though I am not aware that such a compound is made use of, either in chemical or pharmaceutical operations.

The *vinum ferri* of the shops contains the tartrate of potash and iron, in solution, and for all useful purposes, it might be made, at once, by dissolving the triple tartrate in wine. Each pint of the *wine of iron* is said to contain 22 grains of the red

oxyde of iron, but the quantity of oxyde must vary with the proportion of tartar present in the wine. Very dry Sherry will not act upon the iron, until some cream of tartar be added. Some writers direct iron wire to be used, in preference to the filings, and they order Rhenish wine, as being the best solvent, because it contains the most tartar. The solution thus made, constitutes a very pleasant article, and is one of the oldest chalybeate medicines, on the records.

*Gallate of iron* is a compound, of gallic acid 51 parts or one equivalent, and peroxyde of iron 40, or one equivalent; it is, consequently, a pergallate. It may be formed, by adding a solution of gallic acid to a solution of a salt of iron, provided the metal be, in part at least, in the state of peroxyde. A dark colored precipitate is thrown down, which constitutes the basis of *writing ink*, but it is not usual to take the pure gallic acid for this purpose, excepting merely as a chemical experiment.

Many receipts have been published for the preparation of black ink, but the following, by M. Ribaucourt, will give general satisfaction. Boil two ounces of finely chipped logwood, with four ounces of coarsely powdered gall-nuts, in six pounds of rain or river water, till half the liquid is evaporated. Let the decoction, thus made, be strained through a linen cloth, after which, add two ounces of sulphate of iron, (green copperas,) half an ounce of the sulphate of copper, (blue vitriol,) an ounce and a half of gum arabic, and a half ounce of sugar, all in fine powder. Stir the mixture until it acquires a homogeneous appearance, and after allowing it to stand for twelve hours, pour off the clear liquor and bottle it. If the ink be not quite black at this stage of the operation, let the bottles be uncorked for one or two days, in order that oxygen may be absorbed from the air, and peroxyde of iron be completely formed.

In this process, the tannin and gallic acid of the gall-nuts, as well as the gallic acid of the logwood, unite with the oxydes of iron and copper of the salts, forming the dark colored precipitate which is diffused through the liquid, and kept in suspension by the gum and sugar, which give to the whole a suitable consistence. The sulphate of copper gives a deeper shade to the mixed precipitate, but it is not always used. When ink contains copper, it may be easily detected by metallic iron, a portion of which is taken up, on dropping a little ink upon it, and replaced by a coating of metallic copper.

If an aqueous solution of chlorine be poured into some ink, mixed with water, the dark color will instantly disappear. The chlorine destroys the color, by acting on the vegetable matter only; and if a solution of hydrosulphuret of potash be added, sulphuret of iron will be precipitated.

Exp.

To show that the chlorine does not act on the iron, in destroying the color, write on four pieces of paper, with some black ink, and when the writing is quite dry, dip them into the aqueous solution of chlorine, until the letters disappear. Now, as the iron has been unaffected in this operation, the characters may be easily recalled. The papers having been dried, immerse one of them in an infusion of nutgalls, another in a solution of hydrosulphuret of ammonia, and a third in a weak solution of ferrocyanate of potash. The oxyde of iron will be blackened by the infusion of galls, and also by the hydrosulphuret, while the ferrocyanate will make it blue. If the fourth piece of paper be dipped in diluted muriatic acid, (one part acid to twelve of water,) all the iron will be dissolved in a few minutes, a soluble muriate of iron will be formed, and then it will be impossible to recal the characters.

Exp.

Exp.

The reason why ink, dropped on linen, causes an iron mould, as it is called, after being washed, is, that an insoluble pergalate of iron is deposited. As the oxalic and some other acids can readily dissolve the iron thus deposited, they are employed to remove such stains. The oxalate of iron, so formed, is soluble, and of course, can be washed out of the linen.

If paper be written on with a weak infusion of galls, no characters will be seen when it is dry; but if it be dipped into a solution of green copperas, which always contains some peroxyde of iron, the letters will immediately appear.

Exp.

Experiments may be made by writing on paper with a solution of a persalt of iron, and dipping it into a solution of ferrocyanate or sulphocyanate of potash. The former will strike a fine blue color, and the latter a deep red.

Exp.

*Ferrocyanate of iron* is sometimes called *prussiate of iron*, and more commonly, *Prussian blue*. It is also termed the *ferrocyanate of the peroxyde of iron*, because it is formed, by adding a solution of a persalt of iron to a solution of ferrocyanate of potash, as long as any precipitation takes place. The precise proportion of the constituents has not been satisfactorily determined; but it is highly probable, that it contains two equivalents of peroxyde of iron, 80, and one and a half equivalents of ferrocyanic acid, or 162.

If the solution of iron happen to contain but little peroxyde, and to be chiefly a protosalt, the precipitate, formed by adding the ferrocyanate of potash, will be of a light color, at first. But if it be exposed to the air, or if a very little sulphuric acid be added, the state of oxydation will be altered, and the iron be



changed to a persalt. The air in the one case, and the acid in the other, affords an additional dose of oxygen, by which the transition is effected from protoxyde to peroxyde.

The ferrocyanate of the peroxyde of iron is regarded, merely as the basis of the Prussian blue, which is known to contain alumine and peroxyde of iron, together with the subsulphates of one or both of these bases.

Prussian blue was discovered in 1710, and is prepared in large quantities, in various sections of our country. Dried blood, or other animal matters, as the offals of slaughter-houses, are heated to redness in iron vessels, together with an equal weight of pearlash, until the whole acquires a pasty consistence. The fused mass chiefly consists of cyanuret of potassium and carbonate of potash, which must be dissolved in water, and filtered. To the filtered liquor, a solution of two parts of alum and one of the common green copperas, are added, though some manufacturers omit the alum. A dirty greenish precipitate is soon formed, which absorbs oxygen from the air, and changes gradually to the blue shade that is peculiar to the pigment.

The manner in which the elementary principles of the animal matters are acted upon, in connexion with potash and iron, to form the ferrocyanic acid, has been clearly stated, under the head of ferrocyanic acid, and it is necessary to recollect the rationale there given, in order to understand the reaction that occurs, when the filtered liquor above named, is added to the solution of sulphate of iron. The precipitate, which is at first thrown down, is occasioned by the potash, and consists chiefly of alumine and protoxyde of iron. The hydrocyanic acid, formed by heating the animal matters to redness, is converted into ferrocyanic acid, by the reaction of part of the protoxyde of iron contained in the sulphate. In this action, water and cyanuret or cyanide of iron are generated, the latter of which unites with the undecomposed hydrocyanic acid, and gives rise to ferrocyanic (or, more properly, ferrohydrocyanic) acid. This acid then combines with some oxyde of iron, which may be partly in the state of protoxyde, and partly in form of peroxyde. By absorption of oxygen, the whole of the iron is peroxydated, and the compound evinces the characteristic blue tint.

Hence, although the maximum of oxydation is essential to the final production of the deep blue color, yet the presence of the protoxyde of iron is necessary, in the first instance, in order to act on the hydrocyanic acid, to change it to the ferrocyanic, a result which is too feebly accomplished by the peroxyde, to answer any good practical purpose.

When pure, dry Prussian blue is exposed to heat, it soon undergoes a slow combustion, and is eventually resolved into its elementary parts. Its use in preparing cyanogen and hydrocyanic acid, has been already considered.

The ferrocyanate of the peroxyde of iron, has been employed in medical practice, especially in the treatment of intermittents; but it has fallen very much into disuse.

---

## CHAPTER II.

### Of Lead.

THE bluish-gray color of this metal, is familiar to most persons. It has a strong, metallic lustre when fresh cut, but exposure to the air soon tarnishes it. It has a specific gravity of 11.5; its melting point is about  $600^{\circ}$ , and its equivalent or combining number is 104. It is soft, flexible, and elastic. It is both ductile and malleable; the latter property is conspicuous in the large sheets which are made of this metal. In tenacity, it is inferior to all the ductile metals. The compounds of lead are nearly all poisonous.

It has long been a matter of dispute, whether leaden pipes communicate any deleterious property to water. There can be no doubt, that pure water occasions the oxydation of lead, if atmospheric air have free access; and the metal thus exposed, yields minute, brilliant, white crystals of carbonate of lead; the oxygen and carbonic acid being derived from the atmospheric air. But water generally contains saline particles which retard the oxydation of the metal, and some salts appear to prevent that result entirely. Many kinds of spring water, by reason of the saline matter held in solution, do not corrode lead, in the smallest degree; and hence, these waters, even when designed for drinking, may be passed through leaden pipes, and collected, if desirable, in cisterns lined with that metal. So great is the preservative power afforded by these salts, that a four-thousandth part of sulphate of lime is sufficient to prevent the slightest oxydation of the pipes through which the water, containing that salt, flows; a two-thousandth part of muriate of soda, affords equal security, while only a twelve-thousandth part of arseniate of soda, or a twenty-seven-thousandth part of phosphate of soda, is requisite to produce the same effect.

This preservative power depends on the acid, and not on the base of the salt, which the water holds in solution. In general, those salts are most efficacious in securing the water from the injurious effects resulting from the oxydation of lead, whose acids are known to form insoluble salts, when combined with that metal; while those salts whose acids, joined to lead, give soluble compounds, afford less protection, and sometimes are nearly inert. It might be inferred, from these facts, that the acid of the salt, possessed of the strongest preserving power, combines with the lead, and coats it with an insoluble sheath, and so prevents any deleterious result. But this conclusion is not sufficiently supported by experiment, to warrant us in making it positively; and further researches must be made, in order to arrive at a correct view of this subject.

Although pure water, exposed to the air, occasions the oxydation of lead, we are assured, that distilled water, from which the external air is excluded, does not exert such an agency, even in the smallest degree.\*

Lead is not attacked by the muriatic or the vegetable acids. Cold sulphuric acid does not act upon it, and even when boiled, the oxydation is not extensive, and hence the manufacturers of sulphuric acid expose that article to heat in leaden vessels, in order to concentrate it.

It has been ascertained, by actual experiments made on dogs and other animals, that pure metallic lead is wholly free from poisonous properties.

Lead is generally obtained from galena, the native sulphuret of lead. To prepare a small quantity from this ore, reduce 120 grains to a fine powder, mix them with 56 grains of iron filings, and expose the mixture to a bright red heat, for five or ten minutes, in a chauffer with a chimney, or in a furnace. Part of the iron unites with the sulphur of the sulphuret of lead, forming sulphuret of iron, and the metallic lead is melted, and may be poured out when the crucible is removed from the furnace. In the diagram giving a view of the theory of the decomposition, only half the quantity of iron is represented, as 28 parts of iron are sufficient to combine with all the sulphur in 120 of the sulphuret, and an excess is employed solely for the purpose of bringing the sulphur of the sulphuret into more intimate contact with the iron, which is not easily reduced to a minute state of division.

\* See *Annales de Chimie*, vol. lxxi., and *Christison on Poisons*, p. 387.

Before decomposition.		After decomposition.
120 sulphuret of lead.	{ lead 104 sulphur 16	104 lead.
28 iron	28	44 sulphuret of iron.
148	148	148

When lead is required particularly pure, for delicate experiments, it may be obtained by precipitation from a solution of any of its soluble salts. Dissolve an ounce of the crystallised acetate of lead, in 34 measured ounces of water, and put a piece of zinc into the solution, suspending it at the top, by a string fixed to a wire laid across the mouth of the jar or glass containing the solution; part of the zinc is immediately dissolved and the rest coated with metallic lead, which continues to be precipitated till the acetate has been completely decomposed; 34 parts of zinc, (or one equivalent,) being taken up, for every 104 of metallic lead, (one equivalent,) thrown down, so that a solution of the acetate of zinc remains. If the solution is put into a long glass jar, and not agitated, the lead is deposited in an arborescent form, presenting a very beautiful appearance; it is in this manner that the *lead tree*, as it is termed, is commonly prepared. The following diagram shows more precisely the nature of the reaction.

Before decomposition.		After decomposition.
34 zinc	34	93 acetate of zinc.
163 acetate of lead	{ acetic acid 51 oxygen 8 lead 104	104 lead.
197	197	197

There are three oxydes of lead, the composition of which, as ascertained by Berzelius, is as follows:

	Lead.	Oxygen.
Protoxyde	104	+ 8 = 112
Deutoxyde	104	+ 12 = 116
Peroxyde	104	+ 16 = 120

The protoxyde is that which forms the basis of all the salts of lead. It may be prepared, by exposing the film which collects on melted lead, to a heat sufficiently strong to give it a uniform, yellow color, and in this form it is known, as the *massicot* of commerce. When the heat is pushed still further,

so as to fuse this yellow mass, it is changed into a scaly substance, having a mixed color of red and yellow, owing to the formation of some red oxyde, and it is then called *litharge*.

We may also prepare the protoxyde, by exposing the nitrate of lead to heat in a crucible, taking care not to urge the heat too far; otherwise the oxyde will melt, and combine with the silicious matter of the crucible, forming a very fusible glass. In this case, the nitrate of lead is decomposed, its acid being driven off, and the protoxyde remaining.

The protoxyde may be made, likewise, by precipitation. If we add a solution of potash, soda, or ammonia, to a solution of acetate or nitrate of lead, the acid will quit the lead, and combine with the alkali, and the protoxyde of lead will settle in the bottom of the vessel, of a white color, owing to incorporation with a portion of water. The protoxyde, thus formed, is sometimes called the *white, hydrated protoxyde*.

The protoxyde of lead unites readily, as we have seen, with silex, and with other earthy matters, giving rise to a transparent, vitrified mass. On this account, it is used in glazing earthenware and porcelain. It enters largely into the composition of flint glass, which is thus rendered more fusible, clear, and uniform, than green glass. It is also added to wines, to remove their acidity. For the purpose of detecting this poisonous matter, a portion of the suspected liquor should be evaporated to a thick fluid, charcoal added, and the mass be calcined in a crucible. In the space of an hour, metallic points will be formed, consisting of lead surrounded by the yellow protoxyde.\*

But we can also detect the presence of lead, by the following process. Pass sulphuretted hydrogen gas freely into the suspected liquor, or add to it a drachm or two of the hydrosulphuret of ammonia. A black precipitate of sulphuret of lead will be formed, which should be collected on a filter and washed with pure water. Then digest it in nitric acid, diluted with twice its weight of water, until the dark color of the sulphuret disappears. A nitrate of lead is thus formed, which should be evaporated, and further heated, to expel the excess of nitric acid. The residue is then to be dissolved in a small quantity of cold water, and a drop or two of a solution of hydriodate of potash being added, a yellow precipitate of iodide of lead is instantly formed.†

*Litharge*, called also the semi-vitrified oxyde of lead, is employed in the arts, especially to give paint a drying quality. It

\* Paris's Pharmacologia.

† Turner's Chemistry.



does this, by imparting to the oil its oxygen. It is also used in pharmacy, to form *lead plasters*, &c.

The *deutoxyde of lead* or the *red oxyde*, is the minium or red lead of commerce, employed so largely as a pigment, and also in the manufacture of flint glass. It is formed by heating the protoxyde in open vessels, while a current of air is made to play over its surface.

This oxyde does not unite with acids. When heated to redness, it yields pure oxygen gas, and is brought back to the state of protoxyde.

The *peroxyde of lead*, which is of a brown color, may be made by digesting the deutoxyde in nitric acid. The deutoxyde is instantly resolved into protoxyde and peroxyde; the former unites with the acid, and the latter remains as an insoluble powder. But it may be prepared, with more facility, by passing a stream of chlorine through a solution of the acetate of lead. In this process, water is decomposed; its hydrogen joins the chlorine to form muriatic acid, and its oxygen goes to the protoxyde of lead, and changes it into the peroxyde. If this oxyde be exposed to a red heat, decomposition ensues, oxygen gas is evolved, and protoxyde of lead remains.

To show the reduction of metallic lead, we mix 200 or 300 grains of any of the oxydes of lead, with a ninth part of its weight of charcoal, and expose the mixture to a good, red heat in a furnace, or open fire. The carbon combines with the oxygen of the oxyde, forming carbonic oxyde, or carbonic acid gas, and metallic lead remains in the crucible. Exp.

*Chloride of Lead.* This is a compound of one equivalent of chlorine, 36, and one equivalent of lead, 104. It is sometimes called *horn lead*, and when in solution, is regarded as a *muriate of the protoxyde of lead*. It may be formed, though rather tardily, by the action of chlorine gas on thin plates of lead; but is more easily obtained by adding muriatic acid, or a solution of sea-salt to the acetate or nitrate of lead, dissolved in water. The chloride dissolves partially in hot water, especially if a little muriatic acid be added. As the solution cools, small acicular crystals of a white color are deposited. At a heat below redness, it is fusible, and on cooling, a semi-transparent horny mass presents itself.

The *mineral* or *patent yellow* of the painters, is a compound of chloride and protoxyde of lead. It is usually prepared by the action of moistened sea-salt on litharge, by which means a portion of the protoxyde is changed into the chloride. After this, the mixture is exposed to heat, in order to fuse it. The soda of the sea-salt is liberated and changed into a carbonate of soda, by absorbing carbonic acid from the atmosphere.

*Iodide of lead* is formed, as already hinted at, by mixing a solution of hydriodate of potash, or the hydriodic acid, with the acetate or nitrate of lead, dissolved in water. The nature of the reaction is shown in the following diagram.

173 or 1 equiv. of hydriod. potash	{ iodine 124 hydrogen 1 potash 48	9 water. 102 nitr. potash.
166 or 1 equiv. of nitrate of lead	{ nitric ac. 54 oxygen 8 lead 104	228 iodide of lead.
<hr/> 339 <hr/>	<hr/> 339 <hr/>	<hr/> 339 <hr/>

In this process, the iodide of lead is precipitated, of a rich yellow color, and may be separated by throwing the whole on a filter. The nitrate of potash, held in solution by the water, formed by the union of the oxygen and hydrogen, passes the filter, while the iodide remains.

If boiling water be poured on the iodide of lead, a colorless solution is the result; this, on cooling, deposits yellow, crystalline scales of a brilliant lustre, which have been employed as a pigment.

*Sulphuret of Lead.* We have already noticed the native sulphuret, or *galena*, and have now to say, that an artificial sulphuret may be formed by heating a mixture of sulphur and lead, or by passing a stream of sulphuretted hydrogen through a salt of lead, dissolved in water. It consists of one equivalent of sulphur, 16, and one equivalent of lead, 104.

We shall next notice the *salts of lead*, and before we examine them particularly, a few general remarks may not be improper.

Salts of lead, in solution, are easily detected by a stream of sulphuretted hydrogen gas, applied in the usual way; the sulphur of the gas combining with the metallic lead, and forming a dark colored sulphuret which is precipitated, while the hydrogen unites with the oxygen of the oxyde, and forms water. Instead of passing a stream of the gas through the solution, many prefer adding a solution of the carbonate of soda or potash in the first place; the carbonic acid of the carbonate unites with the oxyde of lead, and forms a white colored precipitate, (carbonate of lead,) and on shaking it with sulphuretted hydrogen water, it immediately becomes black. A solution of the hydrosulphuret of ammonia may be used, instead of sulphuretted hydrogen gas.

Exposed to heat before the blowpipe and on charcoal, most of its salts readily give a globule of metallic lead.

Solutions of the soluble salts of lead are transparent and colorless, and give copious precipitates of a white color with alkalis and alkaline carbonates, consisting of oxyde and carbonate of lead. Sulphuric acid and solutions of the sulphates, throw down a white precipitate of the sulphate of lead; muriatic acid and the solutions of muriates, give a white precipitate of chloride of lead, the chlorine of the acid uniting with the metallic base of the oxyde, while its hydrogen combines with the oxygen; hydriodate of potash gives a rich yellow colored precipitate, which is composed of iodine and metallic lead, a reaction taking place similar to what has just been explained with respect to muriatic acid and oxyde of lead.

When silver, bismuth, or mercury is suspected to be present in the same solutions with lead, the black precipitate thrown down by sulphuretted hydrogen cannot be relied on, as an indication of the presence of lead, as these metals also give a dark precipitate with this gas.

*Nitrate of lead* is composed of one equivalent of each of its constituents, viz. 54 of nitric acid, and 112 of protoxyde of lead. We form it, by digesting metallic lead in nitric acid, diluted with seven parts of water, applying heat as long as effervescence continues, and evaporating the solution, after it has been filtered, until a pellicle appears on the surface. Then it is set aside, to crystallise. The crystals are anhydrous, and deflagrate when heated with inflammable matter.

Berzelius procured another compound of nitric acid and lead, which he called a *di-nitrate*, because it contained two equivalents of protoxyde of lead to one of acid.

*Sulphate of lead* is formed, whenever sulphuric acid or a solution of a sulphate is added to a solution of a salt of lead. It is very insoluble and entirely inert. Hence, sulphate of magnesia is given to patients who have swallowed an overdose of the sugar of lead; the double decomposition that occurs, occasions the soluble and poisonous salt of lead to be changed into one that is insoluble and incapable of exerting any deleterious agency.

The sulphate is composed of 40 acid, or one equivalent, and 112 of protoxyde of lead. We may ascertain the quantity of lead, in any solution of a salt of lead, by precipitating it in the form of sulphate of lead, and carefully drying and weighing the precipitate. Every 152 parts of the precipitate, contain exactly 104 of metallic lead.

*Phosphate of lead* may be obtained by adding a solution of phosphate of soda to a solution of acetate of lead. The reaction that ensues gives rise to two new compounds, viz. acetate of

soda and phosphate of lead; the former is held in solution, while the latter, being wholly insoluble, is precipitated.

*Carbonate of lead* has been known by the names of *cerusse* or *white lead*, *subacetate of lead*, &c.

This compound occurs as a native production, but it may be formed by the double decomposition of an alkaline carbonate, and the nitrate or acetate of lead. It is prepared, for commercial purposes, by exposing coils of thin sheet lead to the vapors of vinegar. The acid fumes oxydise the metal, and the oxyde thus formed absorbs carbonic acid from the atmosphere, and so becomes a carbonate of lead.

This carbonate is insoluble in water, but dissolves in pure potash. In addition to its extensive use as a pigment, it is also employed in medical practice for various purposes, and in pharmacy to make the sugar of lead.

Carbonate of lead is often adulterated with chalk or whiting, both of which are carbonates of lime. This admixture is more generally practised, in reference to its use as a pigment, and may be detected by first digesting it in cold acetic acid, and then adding oxalate of ammonia. An insoluble oxalate of lime is thus formed, which denotes the nature of the fraud.

To show the formation of the carbonate of lead, by precipitation, we may add an ounce or two of a solution of carbo-

Exp. *Exp.* *Carbonate of potash*, to the same quantity of a solution of the nitrate or acetate of lead. The carbonate of lead will be precipitated and nitrate of potash be held in solution. The manufacturers of white lead, and also painters who use the same pigment, are very liable to the attacks of a disease known by the name of *colicapictonum*. As this disease is frequently fatal, the reader will be gratified by the perusal of several cases that terminated favorably, on examining the *Lancet*, vol. xvii.

*Acetate of lead* is a very important preparation, in a medicinal point of view. It is composed of 51 parts of acetic acid, or one equivalent, and 112 parts of protoxyde of lead, or one equivalent; and in the crystallised state, there are 27 additional parts or three equivalents of water.

This salt is known, commonly, as the *sugar of lead*, and has been so called on account of its sweetish taste and its general resemblance to white sugar. It is also called *acetated cerusse* and *supraacetate of lead*.

The acetate of lead is made, by dissolving either the carbonate or protoxyde of lead in distilled vinegar, and boiling the solution. This operation should be continued, as long as the acid will take up any portion of the lead; the solution should then be filtered through paper, and evaporated till a pellicle

appears on its surface; crystals of the acetate are deposited as the liquid cools. These are soluble in four parts of water, and the solution has a sweet but styptic taste.

The term *superacetate* was given, on the supposition that the solution of the salt had acid properties, so far, at least, as to redden litmus paper. In this, however, there is a deception, since a solution in water, entirely free from carbonic acid, gives no such indication, in the slightest degree. When sugar of lead is dissolved in ordinary water, the carbonic acid present detaches a portion of the protoxyde of lead, and a carbonate of lead is precipitated. Under these circumstances, the acetic acid must be in excess, for a portion has been entirely separated from its connexion with a portion of protoxyde. This acetic acid joins the undecomposed acetate, and hence we have a bi-acetate of lead, or an acetate with excess of acid. As carbonic acid does not act upon this compound, we have only to add a little acetic acid to the common acetate of lead, to prevent any decomposition by the agency of carbonic acid. The use of this salt, in forming the acetate of alumine, has been already noticed. It is also employed in medical and surgical practice, and in chemistry, as a reagent.

The *incompatibles* of this salt demand particular attention, as mistakes are often committed, arising from ignorance on this subject. The *alkalis*, *alkaline earths*, and their *carbonates*, *most of the acids*, *alum*, *borax*, the *sulphates* and *muriates*, *soaps*, all the *sulphurets*, *ammoniated* and *tartarised iron*, *tartar emetic*, *undistilled water*, and solutions containing *carbonic acid*, are all incompatible.

Some practitioners make use of mixtures of acetate of lead and alum; others, of acetate of lead and Epsom salt. But both prescriptions are chemically incompatible, and must defeat their own intention. Alum and Epsom salts are both sulphates; the former is added to increase the astringency of sugar of lead, and the latter to prevent constipation of the bowels. But in both cases, an insoluble and inert sulphate of lead is formed by double decomposition.

The same objection applies to the mixture of acetate of lead and sulphate of zinc, in solution, if the prescriber expects to derive any advantage from the sugar of lead, as such. It is true, that a very valuable article, viz. the acetate of zinc is formed, by the reaction of the two compounds, but certainly, the *ignorant* compiler of prescriptions is not entitled to credit for the result. The following diagram shows the nature of the decomposition that results, from combining sugar of lead with a sulphate.



Sulphate	{ sulphuric acid	40		
of zinc	{ oxyde of zinc	42		
Acetate	{ acetic acid	51		93 acetate of zinc.
of lead	{ oxyde of lead	112		152 sulphate of lead.
		<hr/>		<hr/>
		245		245
		<hr/>		<hr/>

The formation of an inert sulphate of lead, by mixing the various soluble sulphates with sugar of lead, has enabled us to counteract with success the pernicious tendency of an overdose of this article. Hence, the best remedy or antidote for the acetate of lead, is the sulphate of magnesia, given freely in solution. A part of this sulphate decomposes the acetate, forming an insoluble sulphate of lead, while the residue, performing its office as a cathartic, carries the insoluble compound out of the alimentary canal.\*

It would be improper to give the diluted sulphuric acid (elixir of vitriol) very soon after the exhibition of a powder or pill of sugar of lead, for the reason already given; and therefore, as a general rule, all drinks are improper, immediately after the use of the acetate of lead, excepting cold water, or vinegar and water.

The *subacetate of lead* is prepared by boiling the protoxyde of lead, or litharge reduced to a fine powder, in diluted acetic acid; or a solution of the acetate may be used, which will afford a similar solution with the oxyde of lead more speedily; 189 grains of the crystallised acetate may be taken with 112 of the protoxyde of lead, and about nine ounces of water, previously boiled, to expel the carbonic acid it may contain. It has a considerable resemblance to the acetate, in all its leading chemical relations, but is more easily decomposed, does not so readily crystallise, and is less soluble in water.

This compound is sometimes called a *di-acetate*, because it is supposed to contain two equivalents of protoxyde of lead, to one of acetic acid. But, as analysis seems to make it probable, that three equivalents of protoxyde, or 336, and one of acetic acid, or 51, make up its composition, it is also called a *tri-acetate* of lead. I prefer the term *subacetate*, as being more commonly known, and sufficiently well understood. It is the same article that is described in some of the books, as the *extract of lead*.

\* The reader may find a very satisfactory case of poisoning by the sugar of lead successfully treated, on the plan here recommended, in vol. xvii. of the *Lancet*, page 292.

## CHAPTER III.

## Of Copper.

THE native sulphuret of copper (copper pyrites) yields the greater part of the metal of commerce, as it is not found in the native state, in large quantities. The peculiar red color of metallic copper, distinguishes it from all the metals, excepting titanium, and its lustre may be much increased, by polishing. The specific gravity of pure copper is 8.8, and this is augmented by hammering. The equivalent of this metal is 64, and it melts at a bright red heat.

In a perfectly dry atmosphere, copper is scarcely tarnished; but when exposed to a moist air, it absorbs oxygen, and the oxyde thus made, combines with the carbonic acid of the air, and a carbonate of copper is the result. Exposed to a red heat, the metal absorbs oxygen freely, and is covered with black scales of peroxyde. It is acted on, very feebly, by muriatic and sulphuric acids, and is not affected at all by the vegetable acids, if atmospheric air be excluded. Agreeably to Professor Woodhouse, a piece of perfectly pure copper remained in a vessel of strong nitric acid, without experiencing any alteration. If the metal be in a divided state, as in form of filings, this acid acts with great violence.

If we desire to collect small portions of very pure copper, or to have it in a state of minute division, we immerse a piece of clean zinc in a solution of blue vitriol, (sulphate of copper). The zinc takes the oxygen from the oxyde of copper, and is itself, in part, converted into an oxyde, which combines with the sulphuric acid of the blue vitriol, giving rise to sulphate of zinc, while the copper is precipitated on the zinc, in the metallic form. In place of zinc, we may use iron, and in either case, the slip of metal should be left in the cupreous solution, until its blue color has disappeared.

When a piece of copper is wanted, that has a perfectly bright, metallic surface, it must be exposed to a red heat for the space of ten minutes, and then suddenly quenched in cold water. The copper is speedily oxydated in the fire, and scales of red-colored oxyde are formed on its surface. On dipping the metal into water, contraction ensues, and the oxyde being thrown off, exposes a bright metallic surface.

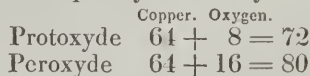
Of the uses of metallic copper, in the arts, we need say nothing here, as these are well known. It is also employed in a vari-

ety of chemical operations, and for the construction of galvanic batteries.

Many facts are on record, to show that metallic copper is not poisonous in the human system.\* The single circumstance of copper filings having been a favorite remedy in the treatment of rheumatism, is abundant evidence on this point. It was given in drachm doses, with perfect impunity.

This feature in the history of copper is the more remarkable, as substances come in contact with it in the stomach, which out of the body would convert it into a destructive agent, as the operations of cookery have often evinced. It is affirmed, that the copper of kitchen utensils cannot be dissolved while tin is present, and hence the utility of tinning copper vessels. So, also, it is stated, that untinned coppers are less likely to do harm, when pewter spoons are used, for stirring the articles cooked in them, than when silver ones are employed; the pewter spoons contain tin, which acts on the same principle as the tin coating of copper utensils. Mr Proust has shown, that the lead, in the mixture of tin and lead, employed to line a variety of kitchen articles, cannot be a source of danger; for so long as the tin is present, the lead cannot be oxydated by any vegetable acid. So much stronger is the affinity of tin for oxygen, than that of lead, that if the latter could possibly unite with a particle of oxygen, the tin would instantly carry it off. Now this may be explained by the idea of the formation of a galvanic circle, but it would seem, that the commonly received notions of elective attraction, are sufficiently explanatory for all useful ends. The cases of copper mixed with tin, and of tin mixed with lead, admit of one and the same explanation; and the plain inference is, that before copper can become poisonous, it must be converted into an oxyde. We shall therefore proceed to consider the compounds, formed by the union of copper and oxygen.

There are but two oxydes of copper, recognised by chemists, viz. the protoxyde and peroxyde. They are thus constituted.



The *protoxyde* is of a red color, and is found in the Cornwall, (England,) mines, in very fine crystals. We prepare it, artificially, by heating to redness, in a close vessel 64 parts or one equivalent of metallic copper, and 80 parts or one equivalent of the peroxyde. The latter is decomposed, yielding one

\* See Paris and Christison.

equivalent of oxygen, which combines with the metallic copper, and reduces the whole mass to the state of protoxyde. It may also be formed, by boiling a solution of verdigris (acetate of copper) with sugar. The peroxyde in the verdigris, is partly deoxydised, and the protoxyde subsides in the form of a red powder.

If a solution of potash be mixed with a solution of protomuriate of copper, the muriatic acid of the latter will unite with the potash, forming a soluble muriate, while the protoxyde of copper is thrown down, combined with a portion of water.

This protoxyde combines with the muriatic, sulphuric, and some other acids, forming salts, which are generally colorless, and which attract oxygen from the atmosphere, by which they are changed into persalts.

The *peroxyde of copper* or the *copper black*, of mineralogists, is sometimes found native, and results from the spontaneous oxydation of the ores of copper. It may be formed, by calcining metallic copper, by precipitation from the persalts of copper, and by means of pure potash, by heating the nitrate of copper to redness. In the latter case, all the nitric acid is expelled, and nothing remains but the peroxyde, quite free from water, and of a very dark color.

Heat alone will not reduce this peroxyde, but if it be exposed in a crucible to a bright red heat, mixed with a seventh part of its weight of charcoal, carbonic oxyde will be disengaged, and a button of metallic copper remain in the bottom of the crucible.

The peroxyde is insoluble in water, and does not affect the vegetable blue colors. It combines with nearly all the acids, and its salts have a green or blue tint. It is precipitated by pure potash, as a blue hydrate, which is made black by boiling, the hydrate being decomposed at that temperature.

If we put some peroxyde of copper into a small bottle, nearly full of water of ammonia, and shake it frequently, part of the oxyde will be dissolved, and the solution assume a rich, blue color. If a quantity of copper filings be added to the solution, and the bottle be well closed, so as to exclude the air, one portion of the metallic copper will combine with Exp. part of the oxygen in the peroxyde, and thus give rise to the protoxyde; and as the peroxyde has now lost one portion of its oxygen, it is also changed into the protoxyde. The changes thus effected, gradually destroy the color of the solution, which in a few days is quite clear, consisting of water of ammonia and protoxyde of copper. If we now open the cork, so as to give the air free admission, the blue color will return, owing to the

absorption of oxygen, by which the protoxyde of copper is changed into the peroxyde.

Put some metallic copper into water of ammonia, and allow them to remain mixed together for some hours; the copper acquires oxygen from the water, and a small portion of  
 Exp. oxyde is formed, which is speedily dissolved by the ammonia, and the usual blue color appears.

Copper is easily detected, when in solution, by the dark, reddish-brown precipitate which is thrown down by solutions of the ferrocyanate of potash and hydrosulphuret of ammonia. Ammonia, added to a solution of any of its salts, unites with the acid and precipitates oxyde of copper, which may be redissolved by an additional quantity of ammonia, the solution acquiring the characteristic blue color.

The oxyde of copper is thrown down, of a yellowish-white color, by albumen, and Orfila has proved, that this compound is inert, so that albumen (white of eggs) is an antidote to poisoning by copper.

*Chlorides of Copper.* Of these, there are two varieties, viz. the *chloride* and *bi-chloride*, called *proto* and *perchloride*. The first contains 64 parts, or one equivalent of copper, and 36, or one equivalent of chlorine, while the latter contains the same quantity of copper, and twice as much chlorine.

If some Dutch gold leaf (composed of copper and zinc) be placed in a bottle of chlorine gas, combustion will take place, with emission of a red light, and chloride of copper, with some chloride of zinc, will be formed.

The chloride of copper may also be made, by heating copper filings with twice their weight of bi-chloride of mercury. The copper takes one equivalent of chlorine from this bi-chloride, and the product consists of chloride of copper and chloride of mercury. The article thus prepared, has some resemblance to common resin, and was called by Mr Boyle, who first made it, the *resin of copper*. Proust formed it, by the action of protomuriate of tin on permuriate of copper, and he called it the *white muriate of copper*. The variety in the color of this chloride, depends on the mode of preparation, and hence we have it white, yellow, or dark brown.

The *bi-chloride* is obtained, by exposing a solution of the permuriate of copper, to a temperature not over 400°, and keeping up the heat till it is evaporated to dryness. It is a yellow, pulverulent substance, which deliquesces on exposure to the air, and is changed by the action of water into a permuriate.

*Sulphuret of Copper.* There are two varieties of this article. The one is a sulphuret, and consists of one equivalent of sulphur, or 16, and one of copper, or 64; the other is a bi-sulphuret,



and contains twice as much sulphur as the simple sulphuret, and precisely the same quantity of copper.

The sulphuret of copper is found in nature, under the appellation of *copper glance*. It may be formed, artificially, by heating a mixture of one part of sulphur and three of copper filings, in a glass flask. A vivid glow of light attends the combination, and of course, there is much heat evolved.

The bi-sulphuret may be formed, by the action of sulphuretted hydrogen gas on a persalt of copper. The hydrogen, of the sulphuretted hydrogen, combines with the oxygen of the oxyde, in the persalt, and the liberated metal joins the sulphur of the sulphuretted hydrogen, whence results the bi-sulphuret. To show this by a diagram, we suppose two equivalents of sulphuretted hydrogen to act on one equivalent of peroxyde of copper.

31 or 2 equivalents of sulphuretted hydrogen	{ sulphur 16		
	{ sulphur 16		
	{ hydrogen 1	-----	9 water.
	{ hydrogen 1	-----	9 water.
80 or 1 equivalent of peroxyde of copper	{ oxygen 8		
	{ oxygen 8		
	{ copper 64	-----	96 bi-sulphu. copper.
114	114	114	

When native sulphuret of copper is roasted, in a situation having a free exposure to the air, both its elements are oxydated to the maximum, giving rise to sulphuric acid and peroxyde of copper, which unite and form the *sulphate of copper* or *blue vitriol*.

The composition of sulphate of copper is two equivalents of sulphuric acid, or 80 parts, to one of peroxyde of copper, also 80; and in the state of crystals, ten equivalents of water must be added, to give its entire equivalent of 250. It may be prepared, also, by the direct union of peroxyde of copper and dilute sulphuric acid. Its crystals are always of a deep blue color, soluble in four parts of cold and in two of boiling water; the solution reddens litmus paper.

It is evident from the composition of this salt, as now given, that it is correctly called a *bi-sulphate*.\*

A real sulphate, composed of one equivalent of acid and one of peroxyde of copper, may be formed, by adding pure potash to a solution of the bi-sulphate of copper, in such quantity as barely to neutralise the excess of acid; a pale bluish-green precipitate falls, which has been erroneously called a *subsulphate*, but which is, in the strictest sense, a sulphate.

\* Sometimes styled an *oxysulphate*.

Blue vitriol, called sometimes *blue copperas*, is used in the arts, and also in medicine. Its watery solution has a harsh, acrid, and styptic taste. The crystals effloresce in a slight degree, and if sulphuric acid be poured on them, there is no effervescence, and this circumstance serves to distinguish the article from the purer specimens of verdigris, with which it has been occasionally confounded. It is entirely insoluble in alcohol. The *incompatibles* of this salt are, all the *alkalis and their carbonates*, borax, acetate of ammonia, tartrate of potash, muriate of lime, nitrate of silver, sugar of lead, corrosive sublimate, all astringent vegetable infusions and tinctures.

If we drop liquid ammonia into a watery solution of blue vitriol, or bi-sulphate of copper, until the precipitate at first formed is nearly all dissolved, we shall have a triple compound, Exp. consisting of sulphuric acid, peroxyde of copper, and ammonia, and hence called sulphate of copper and ammonia. If the dark blue solution, thus made, be concentrated by evaporation, crystals will be deposited on adding alcohol, which takes up the water from the salt, by superior affinity. It may also be made, by rubbing two parts of blue vitriol with three of common carbonate of ammonia, in a mortar, till they become quite moist, and the mixture has a very rich blue color. It should then be dried, with a very gentle heat, and placed in well stoppered bottles. In this process, the ammonia of part of the carbonate, unites with the sulphuric acid and oxyde of copper in the sulphate, and carbonic acid is evolved with effervescence, the water of crystallisation in the sulphate being set at liberty, and rendering it quite moist. The caution about employing a gentle heat to dry the mass, arises from the great volatility of the ammonia, and hence, if the mixture be exposed to the open air, it soon loses its ammonia, and nothing remains but dry sulphate of copper.

This compound has been called *ammoniated copper*, *ammoniuret of copper*, *cuprum ammoniacum*, &c. but I believe the term *sulphate of copper and ammonia*, to be most in accordance with the spirit of modern nomenclature, and with fact.

The taste of this article, in solution, is highly styptic and metalline, and its odor is decidedly ammoniacal, if kept with care. One scruple of the salt dissolves in an ounce of water. It was formerly much used in medical practice, but has been pretty generally laid aside.

*Carbonate of Copper.* This salt is found in nature, as the *malachite* of mineralogists. It is a compound of carbonic acid and the peroxyde of copper, one equivalent of each, and its combining number is 102. It may be formed, by the double

decomposition of carbonate of potash and bi-sulphate of copper, both in solution. The alkaline solution should be added, as long as any precipitation takes place. It may also be made, by exposing metallic copper to air and moisture. The metal is first oxydated, and the oxyde absorbs carbonic acid from the air.

An impure carbonate is formed by the double decomposition of chalk and nitrate of copper. It is used as a blue pigment, and is known by the name of *verditer*.

The pure carbonate is of a greenish-blue color, and dissolves with effervescence in diluted sulphuric, nitric, muriatic, and acetic acids.\*

*Nitrate of Copper.* To prepare this salt, add six parts of copper filings or clippings to four of nitric acid, previously diluted with twice its bulk of water, and digest the mixture for a short time, till effervescence ceases. Then pour off the clear solution, and evaporate it till a pellicle appears on the surface, when it may be set aside to crystallise. During the solution of the copper in the acid, copious irritating fumes are evolved, which ought to be carefully avoided; they are formed by the deutoxyde of nitrogen that is disengaged, attracting oxygen from the air, and being converted into nitrous acid. The metal is oxydated at the expense of a portion of the acid which is decomposed, and the peroxyde thus formed unites with two equivalents of undecomposed acid, and thus produces the nitrate, or rather the *bi-nitrate*. It is composed of 108, or two equivalents of acid, and 80, or one equivalent of peroxyde. It is a very deliquescent salt, and must therefore be kept in close vessels. It has been employed in the London hospitals, very extensively, as an escharotic. The crystals of this salt are of a deep blue color, and of the prismatic form.

If the bi-nitrate be exposed to heat, a green, insoluble salt is procured, which has been called a *subsalt*, but as it consists of one equivalent of acid and one of peroxyde, it is strictly a nitrate. When heated to redness, it yields pure peroxyde of copper.

The experiment usually performed with tin-foil and the bi-nitrate, shows the necessity of water or fluidity to chemical action, and also, the agency of tin in effecting the decomposition of the salt.

Spread a drachm or two of the bi-nitrate, reduced to powder, on a piece of tin-foil several inches square, fold it up quickly, and wrap the whole closely in a piece of stout paper. The paper soon becomes very hot, and cannot be held in the Exp.

\* For an analysis of the *malachite* of Pennsylvania, see a paper by the author, in the *Memoirs of the Columbian Chemical Society*.

hand, owing to the violence of chemical action. The tin decomposes the salt, attracting the oxygen from the nitric acid, by which it is converted into an oxyde; a large quantity of nitrous gas is evolved, and frequently sparks are seen, formed by the burning of small portions of tin. But if the salt be not moistened, it may be closely folded up in the tin-foil, and yet no action will be excited.

*Acetate of Copper.* Several compounds of acetic acid and copper are described in the books. We shall notice, particularly, no more than two, viz. the *bi-acetate* and the *subacetate*; in addition to these, Berzelius describes three others, which do not, however, appear to be important, nor generally acknowledged.

If the verdigris of the shops (which is composed of acetic acid, peroxyde of copper, and some impurities) be digested in five or six times its weight of water for half an hour, in a common oil flask, it will be resolved into two salts, viz. a subacetate which is insoluble, and a bi-acetate which remains in solution; they may be easily separated by filtration, and the solution of the bi-acetate gives crystals on evaporation. Every three equivalents of verdigris give one equivalent of each of these salts; the reaction that takes place is represented in the annexed diagram, omitting the water which the verdigris contains, as it does not suffer any material change.

Before decomposition.			After decomposition.	
Verdigris	{	acetic acid 51	182	bi-acet. copper.
		perox. copper 80		
Verdigris	{	acetic acid 51	211	subacet. copper.
		perox. copper 80		
Verdigris	{	acetic acid 51	393	
		perox. copper 80		
<hr/>			<hr/>	
393			393	
<hr/>			<hr/>	

Verdigris is frequently adulterated with chalk or plaster of Paris. To detect either of these substances, all that is necessary is to mix some with water, in a glass, adding a little sulphuric acid, and stirring the mixture with a glass rod. All the verdigris will be dissolved, and a blue colored solution of the sulphate of copper formed, the acetic acid of the verdigris remaining in the liquid; if any sulphate of lime or chalk should have been mixed with it, a white powder will remain at the bottom of the liquid.

*Bi-acetate of copper* is usually prepared, by dissolving carbonate of copper or verdigris in distilled vinegar and evaporating

the solution, till a pellicle appears on its surface. It gives octahedral crystals on cooling, which contain about a ninth part of their weight of water, and are soluble in nearly five parts of boiling, and twenty of cold water.

*Verdigris* may be formed, by exposing metallic copper to the vapors of vinegar; the metal absorbs oxygen from the air, and the oxyde then unites with the acid of the vinegar. It is prepared largely, in France, by covering copper plates with the refuse of the grape, after the juice has been extracted for making wine. The husks afford acetic acid, by reason of the fermentation of their saccharine matter, and the plates becoming oxydated, a coating of the acetate soon forms on their surface.

The color of verdigris varies from a pale green to a blue, and hence it has been confounded with blue vitriol. It is said, by Mr Phillips, to be composed of one equivalent of peroxyde of copper, 80, one of acetic acid, 51, and six of water, 54. When acted on by sulphuric acid, the acetic acid escapes with effervescence, emitting its peculiar odor, but if the same acid be added to the blue vitriol, it causes no change.

*Ærugo aris*, or the *rust of brass*, is one of the names by which verdigris has been known; it was so called, from the supposition that it was formed in the same manner as the rust of iron.

This substance has been swallowed by mistake, and has produced all the usual symptoms of poisoning. Among the antidotes employed, vinegar has had a conspicuous place; but as it converts the verdigris into a soluble acetate or bi-acetate, it is manifestly improper. The experiments of Orfila have shown, conclusively, that there is no real antidote for the poison of the salts of copper, but albumen. Sugar exerts a chemical action on these substances to a certain extent, by reason of which their solubility is lessened, and hence we may alternate, with the white of eggs, the free use of sugared water.

When the verdigris of the shops is exposed to heat, sufficient to expel its water of crystallisation, an efflorescent mass is obtained, which has been much employed as an escharotic in the treatment of indolent ulcers. It is made into an ointment, with simple cerate.

*Muriate of Copper.* There are two varieties of this compound, viz. the *permuriate* and the *protomuriate*.

The permuriate is prepared, by digesting the peroxyde of copper in muriatic acid, diluted with an equal bulk of water; instead of the peroxyde, we may use the carbonate, prepared by precipitation from the carbonate of potash and the bi-sulphate of copper. The solution of the permuriate has a very dark green color.



The *protomuriate* may be obtained, by mixing copper filings with a solution of the permuriate in a bottle, shaking it frequently and excluding the air. In a few days, the decomposition is complete, and the solution becomes transparent and colorless.

The salts of the protoxyde are generally free from color, while those of the peroxyde have a deep blue or green color.

All the salts of copper give a green tint to inflammable matter, in a state of combustion. This may be illustrated, by mix-

Exp. ing ten or twelve grains of permuriate of copper with half an ounce of alcohol, and exposing the mixture to heat over a spirit lamp, inflaming the alcohol as soon as it boils.

Exp. Or we may make a small spirit lamp with alcohol, colored with the permuriate; on kindling the wick, it will burn with a green flame.

In addition to the remarks made on poisoning with verdigris, it is proper to say, that the deleterious action of all the salts of copper, must be counteracted in the same manner.

It is well known, that these compounds are used by confectioners, to give coloring to their articles; and that, recently, some bakers in France have been detected in using blue vitriol in the fermentation of bread, as a substitute for yeast. These scandalous practices cannot be too severely reprobated, and should be banished from society, by the infliction of heavy penalties.

The accidental formation of acetate of copper sometimes occurs in cookery, and we may often detect, in the vinegar of pickles, a very considerable portion of this salt, simply by adding liquid ammonia, which occasions a blue color.

The most delicate and certain mode of ascertaining the presence of copper, in solution, is by freely passing through the suspected fluid, a stream of sulphuretted hydrogen gas. If copper be present, there will be a pale-brown, muddy-looking precipitate, of sulphuret of copper, which on exposure to the air, acquires a darker color. If this precipitate be heated to redness, to expel the sulphur, and then be digested in nitric acid, we shall have the nitrate of copper. If a few drops of ammonia be now added, a deep blue color will be caused in the solution.

## CHAPTER IV.

## Of Zinc.

THE commercial name of zinc is *spelter*, of which large quantities are imported into this country, for various purposes connected with the arts. This is obtained, either from the native carbonate of zinc, called *calamine*, or from the sulphuret, abundantly diffused in the native state, and known by the name of *zinc blende*. When the former article is used, charcoal is combined with it, and a strong heat is employed. The carbonic acid is thus expelled, and the charcoal deoxydates the oxyde of zinc, and so reduces it to the metallic state. The sulphuret is roasted to drive off the sulphur, and as the zinc is oxydated in this process, the addition of charcoal is necessary, precisely as in the other case. It will easily be understood, from these statements, how it happens that zinc, as it is usually extracted from its ores, is never quite pure; but contains charcoal, sulphur, and small particles of other metallic matter. Hence, also, we see how it happens, that hydrogen gas, made from common zinc, is not perfectly pure; and we are able to appreciate the direction given in the books, to subject it to the process of distillation, in order to free it entirely from foreign matters. If the zinc be placed in an earthen retort, connected with a receiver nearly full of water, and a white heat be applied, the metal will be converted into vapor, which, being condensed in the receiver, furnishes pure zinc.

This metal has a strong metallic lustre, especially when newly fractured, and it has a bluish-white color. Its texture is lamellated; its specific gravity is 7; it melts at  $700^{\circ}$ , and its equivalent number is 34. The file exerts a very feeble action on it, thereby indicating its hardness. At low or high temperatures, it is quite brittle; but between  $210^{\circ}$  and  $300^{\circ}$ , it is malleable and ductile, and hence we find it in large, well formed sheets, not thicker than pasteboard. In this state, it is used for galvanic batteries, and for covering houses. At  $680^{\circ}$  it fuses, and if slowly cooled, it crystallises. Exposed to a white heat, in close vessels, it sublimes unchanged.

*Granulated zinc* is the metal in a divided state, and fitted for chemical use. It is made by melting a given quantity of zinc in an iron ladle, and pouring it into a bucket or other vessel, nearly full of water. Although this metal has been classed with the brittle metals, it is not an easy matter to break it into small pieces, by means of a hammer.

We know of but one oxyde of zinc, and that is composed of one equivalent of oxygen, or 8, and one equivalent of zinc, or 34, making the equivalent of the oxyde 42.

There are two modes by which this oxyde may be made; the first is, by the simple action of heat and exposure to the air; and the second is, by adding an alkali to the sulphate of zinc, in which case, the oxyde is precipitated.

Put half an ounce of zinc into a crucible, capable of containing eight or nine ounces of water, and expose it to a good red heat in a furnace or open fire. A crust of oxyde soon gathers on its surface, and in a short time, part of the metallic zinc begins to be volatilised, burning with a rich, bluish-white flame; if it is then removed from the fire, inclining the crucible to one side, and the oxyde detached from the surface as fast as it is formed, a considerable quantity may be obtained; a portion of the oxyde is, at the same time, carried up mechanically by the current of hot air arising from the crucible, and has usually the appearance of very fine wool. If the operation be carried on in a small sheet iron furnace, with a chimney attached to it, the oxyde will be driven about the room, in copious, white, feathery flakes. This is one of the pharmaceutical processes for preparing *oxyde of zinc*.

Various names have been given to this oxyde, as *nihil album*, *philosophers' wool*, *pompholix*, *flowers of zinc*, &c.

If a solution of potash be added to a saturated solution of sulphate of zinc, diluted with fifteen times its bulk of water, a precipitate will fall instantly, and the alkali should be added as long as any precipitate is thrown down. Should the alkali be added, after the formation of the precipitate ceases, it will redissolve it, and thus the oxyde will be lost.

The oxyde of zinc was formerly employed in medical practice, particularly in the disease commonly called St Vitus's dance, but it is seldom used at the present day.\*

If this article be heated to redness, it changes from white to yellow, but the latter color disappears, as the oxyde becomes cold. In all other respects, it is unaffected by the fire. It is insoluble in water, and does not alter the vegetable blue colors. It furnishes a salifiable base, which unites with acids, yielding colorless salts.

Zinc may be distinguished, when in solution, by the yellowish-white precipitate which it gives with solutions of hydrosulphuret of ammonia, hydriodate and ferrocyanate of potash; by the

\* An impure oxyde, called *tutty*, has been used, in form of ointment, in cases of ophthalmia.

precipitation of its oxyde by means of potash; by heating the oxyde at the blowpipe, on charcoal, thus reducing the metal, and causing it to burn with its characteristic flame.

*Chloride of Zinc.* This article is made by vaporating the muriate of zinc to dryness, and then heating it to redness in a glass tube. It is also called *butter of zinc*. It deliquesces on exposure to the air, and is reconverted into a muriate. It is a compound of chlorine 36, or one equivalent, and 34 of zinc, or one equivalent.

*Sulphuret of Zinc.* We have already alluded to this article, under the name of *zinc blende*, as one of the sources whence metallic zinc is obtained. It may be formed, artificially, by heating to redness a mixture of oxyde of zinc and sulphur, by decomposing sulphate of zinc by charcoal, or by drying the white precipitate, obtained on adding hydrosulphuret of ammonia to a salt of zinc. The rationale of the action in the latter case, is shown by the following diagram.

34 or 1 equiv. of hydro-sulphu- ret of ammon.	{ sulphur 16 hydrogen 1 ammonia 17	
		57 sulph. ammonia. 9 water.
82 or 1 equiv. of sulphate of zinc	{ sulph. ac. 40 oxygen 8 zinc 34	50 sulphuret zinc.
116	116	116

*Sulphate of Zinc.* This salt is found in nature, but generally in too impure a state for useful purposes. It may be formed, in crystals, by evaporating the solution that remains after the preparation of hydrogen gas, from zinc, water, and sulphuric acid. For commercial purposes, it is made by roasting the native sulphuret of zinc, in a reverberatory furnace. The heat causes the attraction of cohesion between the particles to be destroyed; and the sulphur, as well as the zinc, is oxydated to the maximum, forming sulphuric acid and zinc, which, united, make the sulphate of zinc. It is composed of 40 parts sulphuric acid, and 42 oxyde of zinc, being one equivalent of each. In the crystallised form, it contains also 63 parts or seven equivalents of water. The crystals dissolve in two parts and a half of cold water, and in rather less of boiling water. The taste of this salt is strongly styptic. Its watery solution reddens vegetable blue colors, although it is strictly a neutral salt. It is insoluble in alcohol.

Sulphate of zinc is known by the names of white vitriol, white copperas, &c. Its *incompatibles* are, the *alkalis*, *earths* and their *carbonates*, the *hydrosulphurets*, *astringent vegetable infusions*, and *milk*.

There is little danger of sulphate of zinc operating as a poison, for its emetic property causes it to be expelled from the stomach, very speedily. On account of this property, it has been introduced into nearly all the empirical remedies for whooping-cough.

*Carbonate of Zinc.* This compound is met with in the native form, as *calamine*; but it is artificially prepared, in a purer state, by the double decomposition of carbonate of potash and sulphate of zinc. If these be mixed in solution, two new compounds will be formed, viz. the sulphate of potash and carbonate of zinc, the former of which remains in solution, while the latter, being insoluble, is precipitated. To insure the success of the operation, the ingredients should be dissolved, separately, in eight or nine times their weight of water, that a sufficient quantity of that fluid may be present, to hold the sulphate of potash in solution. Seventy parts of carbonate of potash are required, for the decomposition of 145 of the crystallised sulphate of zinc.

This carbonate, or as it is more commonly called, *lapis calaminaris*, is a buff-colored substance, with some red particles interspersed among it, and having an earthy aspect. The principal use of this article, is to form the *calamine* or *Turner's cerate*, called also, *epulotic cerate*.

*Nitrate of Zinc.* This salt may be made, by adding nitric acid, diluted with five or six times its weight of water, to metallic zinc, in the granulated state. After the metal is dissolved, the solution should be evaporated until a pellicle appears on its surface. Part of the acid is decomposed in the commencement of the process, and oxygen is thus furnished to oxydate the zinc; the undecomposed acid unites with the oxyde thus formed, and nitrate of zinc is the result.

This salt is composed of one equivalent of each of its constituents, viz. 54 acid and 42 oxyde; but in the crystallised state it contains 54 parts or six equivalents of water.

*Acetate of Zinc.* This salt may be obtained, by digesting metallic zinc or the oxyde in acetic acid, and concentrating the solution by evaporation, when it is deposited in small crystals. It may also be formed, by mixing a solution of the acetate of lead with a solution of the sulphate of zinc; a double decomposition occurs, in the manner represented in the diagram, page 396. The sulphate of lead, being insoluble, is precipitated,



and the acetate of zinc remains in solution. As both salts contain water of crystallisation, in the form they are usually met with, allowance must be made for this, in weighing out the materials. The solution of the acetate, according to the Edinburgh college, is prepared by mixing 60 grains of the sulphate of zinc with 80 of the acetate of lead, dissolving each previously in ten ounces of distilled water, and filtering the liquid afterwards to separate the sulphate of lead. When made in this way, however, it is apt to retain some sulphate of lead, in solution. To avoid this impurity, it may be formed, by suspending metallic zinc in a dilute solution of the acetate of lead, as in making the leaden tree. The vessel should not be disturbed, until all the lead is removed, or rather, reduced to the metallic state. This may be known, by the addition of a few drops of hydrosulphuret of ammonia, or by passing in a stream of sulphuretted hydrogen, either of which will give a pure white precipitate.

*Muriate of zinc* is always formed in the process of making hydrogen gas, if, instead of diluted sulphuric acid, we employ weak muriatic acid, to act upon zinc. The metal is oxydated, by the oxygen furnished by a portion of decomposed water, and the muriatic acid combines with the oxyde so formed, giving rise to the muriate of zinc.

---

## CHAPTER V.

### Of Antimony.

ANTIMONY\* is a metallic substance that has been known for a long time. The exact date of its discovery is not well ascertained, and for our purpose, is of no importance. It is certain, however, that towards the close of the fifteenth century, Basil Valentine published a treatise on this metal, and detailed the first process for its extraction from the ore. Perhaps no metal whatever has experienced a greater fluctuation in public opinion, especially among medical men; at one time its praises have been lauded to the skies; again it has been denounced in the most positive manner. There cannot be a doubt, however,

\* The derivation of this word is obtained from two Greek words, signifying *against a monk*, because, as it is alleged, a monk killed his brother monk, by the use of this medicine.

that it has furnished to the healing art some very important remedial agents, and veterinary medicine, especially, has laid it under contribution, to a large extent. Of these points, we shall speak more fully hereafter.

The alchemists had unfortunately imagined, that this metal would very greatly facilitate their favorite project, the conversion of baser metals into gold, an object which was prosecuted by them with indefatigable zeal. The white color and lustre of antimony, almost equal to that of silver, seemed to promise an easy conversion into gold. I need not say, that their hopes were delusive, and that in the same furnaces in which the baser metals were volatilised, their prospects were also dissipated and lost. Notwithstanding this want of success, the efforts of the alchemists, in this matter, were not wholly in vain, since they were productive of many highly important observations, of which subsequent philosophers have availed themselves, and thus improved our science. Facts are not governed by the circumstances of fluctuation and mutation, which control theories; the former have a durable existence; they are not the sport of every wind, but stand, whether as monuments of folly or wisdom, firm as the rock on which the surges beat in vain.

Although antimony has been known a very considerable length of time, there are still doubts entertained about some points of its history, and especially of its artificial oxydes. This metal exists naturally, in several forms. 1st. In the metallic state, in white masses, having a lamellated tissue, very brittle, and occasionally mixed with arsenic. This mineral is quite rare, and is, perhaps, an oxyde. 2d. In the state of oxyde, in masses of a pearl-white or yellowish color, arising, as some think, from the decomposition of the sulphuret. 3d. In the state of red sulphuret or the native kermes, in form of thin needles, often interlocked, of a bright red color. 4th. In the state of sulphuret, in masses of a steel-gray color, brittle and fusible, having some analogy with the crystallised oxyde of manganese. This latter ore of antimony is the one most frequently met with, and the same that is usually worked to procure the metal itself.

The very great fusibility of the sulphuret of antimony, promotes its separation from its gangue. A very simple apparatus formerly served for this purpose, consisting of two earthen pots, placed the one inside of the other; the upper pot being pierced with holes at the bottom, was filled with the ore coarsely powdered, and a fire was kindled around the pots; as the metal melted, it flowed into the inferior pot. Instead of these pots, crucibles have been employed and placed in a furnace, so that a communication might be preserved, by means of earthen pipes, with receivers properly situated.

The sulphuret of antimony has been improperly called *antimony*, or *crude antimony*, and it is yet known in the shops under the latter name; the metallic antimony has also been called the *regulus of antimony*; all the ancient formula have these terms, and it is well for those who read the recipes of the older writers, to know the grounds of distinction.

In order to extract metallic antimony, we take the pulverised sulphuret, having previously roasted the ore to separate as much as possible of its sulphur. To this end, the ore is heated until it is melted, being frequently stirred during this operation. The sulphur is thus disengaged in form of sulphurous acid, while the metal passes to the state of oxyde, but still combined with a portion of sulphur, which cannot easily be detached. In this state, it is to be mixed with a half part of calcined tartrate of potash, (composed of potash and tartaric acid); the hydrogen and carbon of the tartaric acid combine with the oxygen of the metal, and water and carbonic acid are produced. On the other hand, the sulphur combines with the potash, and gives rise to a sulphuret of potash, which being very fusible and lighter than the metal, covers the latter and protects it from the action of the external air. Thus the metal is reduced and is found at the bottom of the crucible, whence it may be separated when the apparatus is cold.

This process is simple and convenient. Formerly, another was in use, and it has been revived recently. It consists in mixing the ore of antimony with iron filings, and melting the mixture in a crucible of iron, made into a conical shape. The sulphur has a much stronger affinity for iron than the antimony has, and will therefore leave the latter, to join the former. The sulphuret of iron, formed by this decomposition, is lighter than antimony, and occupies the upper part of the crucible, while the antimony falls to the bottom and is there protected from the action of the air, by means of the superincumbent sulphuret of iron.

We may procure the metal of antimony, in a small way, by mixing three parts of the sulphuret, two parts of tartar, and one part of nitre. The mixture is to be placed in a crucible and heated, until the mass is melted completely. On cooling the crucible and breaking it, very brilliant, small pieces of antimony will be seen, which require a heat of about  $750^{\circ}$  for their fusion. The equivalent of metallic antimony, or the *regulus of antimony*, is 44, and its specific gravity, 6.8.

Metallic antimony contains, generally, a little iron and lead, which existed previously in the sulphuret. M. Vauquelin was the first to notice, that when the mixture was melted a long

while with tartar, it contained a quantity of an alloy of potassium, sufficient to decompose a portion of water in which it was thrown, hydrogen being evolved and the water being rendered obviously alkaline. M. Serullas has also noticed, that when the sulphuret of antimony contains arsenic, (which is almost always the case,) the alloy of potassium, resulting from the tartrate of potash, evinces in its decomposition of water, the presence of arsenic, by the disengagement of a portion of arseniuretted hydrogen. By this method, he detected the presence of an atom of arsenic in the *kermes mineral* and *emetic tartar*. In the decomposition of emetic tartar, by means of the crude tartar, M. Serullas obtained an alloy of potassium, so pregnant with this metal, that a single drop of cold water thrown on its surface, produced an explosion.

Metallic antimony is very brilliant, having a silvery lustre, with rather a bluish cast. It is not altered much more than silver, by exposure to the air, its surface being made a little dull. The following experiment illustrates some of the properties of this metal, and may be easily performed.

Expose two or three hundred grains of metallic antimony to a bright red heat, in a crucible placed in a furnace; the blue flame with which it burns will be seen very distinctly, and a large quantity of white fumes is produced, formed by the combination of the antimony with the oxygen of the air. If the melted antimony is poured out of the crucible at this high

Exp. temperature, from a height of ten or twelve feet, and allowed to fall upon a stone floor, it instantly divides into an infinite number of minute globules, all of which run from the part on which it falls, like radii from a common centre, still burning, and leaving a black mark as they roll along, lined on both sides with a white smoke, and producing a very large quantity of fumes.

If the fumes or vapors thus occasioned, be condensed by coming in contact with cold surfaces, small shining needles of silvery whiteness are formed, which have been called the *argentine flowers of antimony*. This preparation was once employed as an emetic.

*Oxydes of Antimony.* Chemists have held various opinions, respecting the number of compounds which antimony forms with oxygen. Thenard contends for six, while Proust admits but two. The more general view is, that there are three oxydes, viz. the protoxyde, deutoxyde, and peroxyde, whose composition is as follows:

	Antimony.	Oxygen.
Protoxyde	44, or one equiv. and	8 = 52
Deutoxyde	44, “ “ “ “	12 = 56
Peroxyde	44, “ “ “ “	16 = 60

The *protoxyde* of antimony may be prepared in several ways; first by the process named, for making the *argentine flowers* which are really a protoxyde, possessing all the properties of such an oxyde, although some chemists consider them as a deutoxyde. By calcining antimony in open vessels, the metal becomes incandescent before the vessels are heated to redness, and it is changed into a grayish-white matter, exactly resembling the protoxyde. The white precipitate, formed by adding water to the chloride of antimony or by pouring water into the muriate or hydrochlorate of that metal, is also a protoxyde, called the *powder of Algaroth* or *mercurius vite*. This is more properly termed, a hydrated protoxyde of antimony.

The protoxyde, prepared by the last mode, always has a dirty white appearance, which changes somewhat to a yellow, by the application of heat. At a dull red heat, it is fusible, being converted into a yellowish mass, which becomes opaque, gray, and crystalline, on cooling. Heated in contact with the air, it volatilises, and the vapors, on cooling, are changed into needle-shaped crystals, that are quite soluble in acids.

The protoxyde is the only oxyde that forms regular salts with acids, and it is the base of the tartar emetic of the shops, which is strictly a tartrate of antimony and potash.

In addition to the pure protoxydes of antimony, there are several compounds, in which the protoxyde is combined with small portions of sulphuret of antimony. In this variety, we include the *glass, liver, and crocus* of antimony.

If one thousand grains of sulphuret of antimony, mixed with the same quantity of nitre, be placed in a paper cone, and a red hot iron wire be applied to the surface of the mixture, a rapid deflagration follows. The experiment Exp. should be made on an earthen dish or iron plate. A large quantity of fumes will be disengaged, and a dark brown substance remain, mixed with some saline matter, which is chiefly the potash of the nitre. The nitric acid of the saltpetre affords oxygen to the sulphuret, by which the sulphur is changed to sulphuric acid, and the antimony into the protoxyde of that metal; the sulphuric acid, at the same time, combining with the base of the nitre. By reducing the whole mass to a fine powder, and washing it repeatedly with boiling water, all the saline matter is dissolved, and the protoxyde is obtained; still mixed, however, with some sulphuret of antimony, which has not



been decomposed. In this state, the Edinburgh college employ it for making tartar emetic. Its former name was the *crocus* of antimony, thus called from its saffron color.

The impure protoxyde, called *glass* of antimony, may be obtained, by exposing the sulphuret, in powder, to heat in the open air. The sulphur, or a part of it, is converted into sulphurous acid, and escapes in the gaseous form. Enough oxygen is absorbed, however, to constitute the protoxyde, which is mixed with a portion of undecomposed sulphuret of antimony. If the mass, thus formed, be exposed to a stronger heat, in a crucible, it is melted into a glass, which has a reddish color, and is perfectly transparent. There is also present, in this glass, a portion of silex, derived from the earthy matter of the crucible.

This glass of antimony is used in the process of the London college for making tartar emetic; and it is still employed by some persons, in preparing the *vinum antimonii*. For this purpose, the Edinburgh Dispensatory gives the following formula.

Vitrified antimony, in powder, one ounce;  
Spanish white wine, one pint and a half.

Digest for twelve days, frequently shaking the vessel, and then filter through paper. It is a fact, however we may account for it, that the same quantity of glass of antimony will serve for repeated additions of wine, for a great length of time. After thirty repeated infusions, it has been found scarcely diminished in weight.

*Deutoxyde of antimony* is prepared in several ways. Berzelius directs us take the powder of metallic antimony and heat it in nitric acid, which should be at first weak, and afterwards, the concentrated acid should be employed. The solution is then evaporated to dryness, until no more vapors are driven off. The residue is a white hydrate of the peroxyde, which by exposure to a red heat, is changed into the deutoxyde. It is composed of 44 antimony (one equivalent) and 12 oxygen (one equivalent and a half).

The deutoxyde may also be made, by exposing the protoxyde to heat and air; it absorbs oxygen sufficient to convert it into the deutoxyde. If the peroxyde, by whatever method it may be formed, be exposed to a red heat, a portion of oxygen is expelled, and deutoxyde remains.

The deutoxyde is white, infusible, insoluble in acids and in water. It combines with alkalis, and has therefore been called an acid. Berzelius styles it *antimonious acid*, and its compounds, *antimonites*.

This acid, or deutoxyde of antimony, may be precipitated from the antimonites, as a hydrate. Under these circumstances, it reddens litmus paper, and is dissolved by muriatic and tartaric acids, but no definite compounds appear to result from the action.

*Peroxyde of Antimony.* This is the tritoxyle of some writers; but as it is the maximum of the oxydation of antimony, as generally acknowledged, the term peroxyde is preferable.

To prepare it, metallic antimony in fine powder, or the pure protoxyde, may be added to hot nitric acid, in a green glass flask or saucer, evaporating the solution to dryness, and heating what remains to the temperature of  $500^{\circ}$  or  $600^{\circ}$ , to expel any water which it may still contain. The temperature must not be increased beyond this, otherwise part of its oxygen will be disengaged; 200 grains of metallic antimony, and an ounce and a half of acid, by measure, will be a sufficient quantity of materials to show the nature of the process, and the appearance of the peroxyde.

It is also procured, by calcining to redness, in a crucible, a mixture of one part of powdered antimony and two parts of nitrate of potash. The mass, thus obtained, is to be mixed with cold water, and the insoluble white powder is to be treated with nitric acid, which separates the potash. The residue is called the *pearly matter of Kerkringius*, or the pure peroxyde of antimony. This oxyde is still more disposed than the deutoxyde, to act the part of an acid, by combining with bases; hence, it has been called the *antimonic acid*. Berzelius uses the terms *stibious* and *stibic acids*, as well as *antimonious* and *antimonic acids*, to express the same things. Hence we find the terms, *stibiated tartar* and *antimoniated tartar*, in some of the older books.

The peroxyde, or *antimonic* or *stibic acid*, shows, in its composition, its relation to the protoxyde, conformably to the law of definite proportions. It is composed of 44 antimony, and 16 oxygen, whereas the protoxyde contains but one equivalent, or 8 of oxygen.

When the peroxyde is recently precipitated, it reddens litmus paper, and may be dissolved in water by means of muriatic or tartaric acids. It does not form any definite compound with these acids, but with the alkalis it forms salts, which have been called *antimoniates*.

The peroxyde, as obtained by the above process, always contains water, and when it is exposed to a heat of  $550^{\circ}$ , its water is dissipated, and a pure, yellow colored peroxyde remains.

Exposed to a red heat, it parts with oxygen, and is reduced to the state of deutoxyde.

The *oxyde of antimony* with *phosphate of lime*, or *pulvis antimonialis*, which was introduced as a substitute for James's powder, is prepared from a mixture of equal weights of hartshorn shavings and sulphuret of antimony. The mixture is exposed to heat in a shallow iron or earthen vessel, till it assumes an ash-gray color, when it may be considered as composed of protoxyde of antimony and phosphate of lime, the latter being derived from the hartshorn shavings, while all the animal matter is burnt away. It is reduced to powder, and put into a crucible, which may be coated with clay, and wrapped round with iron wire, luting on a cover, but leaving a small opening at the side; the crucible is then put into a furnace, and exposed to a white heat for two hours, after which it may be removed.

This is the process which is recommended by the Edinburgh and Dublin colleges; the London directs two parts of hartshorn shavings, to be used along with one of the sulphuret of antimony.

According to the analyses of Pearson, Brande, and Phillips, it is usually composed of peroxyde of antimony and phosphate of lime, the proportion of these ingredients varying in different specimens; but occasionally a small quantity of the protoxyde is found along with the peroxyde, which explains the circumstance that it sometimes proves an active medicine, while in other cases it has been found to be perfectly inert, the peroxyde having little or no action on the animal economy. It has not yet been determined, whether the oxyde is merely mixed with the phosphate of lime, or in some peculiar state of combination. The uncertainty of this antimonial powder, according to a late writer,\* depends entirely on the method of preparation, and particularly, on the degree of heat employed. If the heat be so great as to drive off all the sulphur from the sulphuret of antimony, in a very short time, the metal acquires a grade of oxydation above the protoxyde; or the protoxyde, if really formed, may be volatilised. As the presence of the protoxyde is essential to the medicinal operation of the powder, it is plain, that its absence destroys the powers of the article. These facts prepare us for a correct understanding of the contradictory reports, respecting the efficacy of the *pulvis antimonialis*.

We ascertain the presence of antimony, in solution, by the action of sulphuretted hydrogen gas. It causes an orange colored precipitate, which is the hydrated protosulphuret of antimony. This is soluble in pure potash, and when hot muriatic

\* See Maugham's Manual of Medical Chemistry.

acid is added, there is an evolution of sulphuretted hydrogen gas; the muriatic acid combines with the antimony, and if water be added to it, the *pulvis algarothi*, (called, by some, a submuriate, but which is really a protoxyde,) is thrown down.

*Chloride of Antimony.* Chlorine combines with antimony in two proportions, giving rise to the *chloride* and *bi-chloride*, or the *proto* and *perchloride*.

The chloride may be formed, readily, by the combustion of metallic antimony in chlorine gas. The best mode, because the simplest, that I know of, is to take a glass funnel, and to fit it accurately to the mouth of the bottle containing the chlorine; and place a small cork or paper plug in the open neck of the funnel, to prevent the escape of the gas. Exp. The antimony, in very fine powder, may then be introduced, in very small quantities at a time, first removing the cork or plug. The lower aperture of the funnel should not project more than a half inch into the bottle, because the more chlorine the metal passes through, the more complete will be the combustion. The product is the chloride of antimony, containing 44, or one equivalent of metal, and 36, or one equivalent of chlorine. It may also be made, by distilling a mixture of metallic antimony and three times its weight of corrosive sublimate. The volatile chloride passes over and is condensed in the receiver, while metallic mercury remains in the retort.

This compound is usually called *butter of antimony*, on account of its soft consistence. It melts when exposed to a gentle heat, and exhibits a crystalline texture, on cooling. It deliquesces when exposed to the air, and if water be added to it, the mass is changed into muriatic acid and protoxyde of antimony. The butter of antimony was formerly employed as an escharotic.

The *bi-chloride* is formed, by passing chlorine gas over heated metallic antimony. It contains the same proportion of metal as the chloride, but has double the quantity of chlorine, and hence the prefix *bi*. This bi-chloride is a transparent, volatile liquid, giving out fumes on exposure to the air. When water is added to it, muriatic acid is formed, and also the hydrated peroxyde, which falls to the bottom of the vessel.

It may seem strange, that the addition of water to the *bi-chloride* should convert it into muriatic acid and *peroxyde*, while the same fluid acting on the *chloride*, produces muriatic acid and *protoxyde* of antimony. This difference, however, depends on the addition of twice as much water in the former case as in the latter, and which is necessary to bring about the change.

To make this matter appear perfectly plain, we subjoin the following diagrams.

9 or 1 equiv. of water	{ hydrogen 1	37 muriatic acid.
	{ oxygen 8	
80 or 1 equiv. chlo- ride of antimony	{ chlorine 36	
	{ antimony 44	52 protox. anti.
<hr/> 89	<hr/> 89	<hr/> 89
18 or 2 equiv. of water	{ hydrogen 1	37 muriatic acid.
	{ hydrogen 1	37 muriatic acid.
	{ oxygen 8	
	{ oxygen 8	
116 or 1 equiv. of bi-chloride of antimony	{ chlorine 36	
	{ chlorine 36	
	{ antimony 44	60 perox. of anti.
<hr/> 134	<hr/> 134	<hr/> 134

In addition to the chlorides above named, Dr Thomson has announced a third, about which, however, there is not a perfect agreement, among chemists. He says, it contains two equivalents of metal, to one of chlorine, and hence he has called it a *di-chloride*.

*Sulphurets of Antimony.* There is no subject in chemistry, in which so much confusion has existed, as in relation to the sulphurets of antimony, and especially, in regard to what have been called the *precipitated* sulphurets. This difficulty is, perhaps, dependent on variety of manipulation, which from the same materials and quantities, occasions diversity in the results. For all useful purposes, we might restrict ourselves to but one real sulphuret of antimony, corresponding in point of equivalent numbers, with the native sulphuret. But Mr Rose, and some others, name a *sesqui-sulphuret* and a *bi-sulphuret* of this metal.

The sulphuret is composed of 16 sulphur, one equivalent, and 44 antimony, one equivalent. The *bi-sulphuret* has the same quantity of antimony, and two equivalents of sulphur.

The *sulphuret* occurs native, and is the ore of antimony from which the metal is usually procured; it is separated from the stony matter with which it is mixed, by exposing the ore to heat in a crucible with an aperture in the bottom, another crucible being placed below it to receive the melted metal as it flows out. It is better always to purchase it in the form in which it is taken out of the crucible, the pounded sulphuret being frequently adulterated with a large quantity of earthy matter; I



have met with it containing at least a third part of its weight of earthy substances, the greater part of which was silica.

By transmitting a stream of sulphuretted hydrogen through a solution of tartar emetic, or of any other solutions of antimony, a copious precipitate, of a deep reddish-brown color is thrown down, which is composed of sulphuret of antimony and water, the hydrogen of the sulphuretted hydrogen combining with the oxygen of the oxyde, and the sulphur with the metallic antimony.

To prepare the *precipitated sulphuret* of the Edinburgh and London colleges, the common sulphuret is boiled with a solution of caustic potash for an hour or two, in an iron vessel over a common fire, adding water from time to time, that there may be the same measure of liquid on moving it from the fire, as at first. It is filtered immediately through a double linen cloth, and sulphuric acid, previously diluted with six or seven parts of water, dropped into the filtered liquid (while still warm) as long as any precipitation takes place.

In this process, a portion of water is decomposed, the hydrogen combining with the sulphur of the sulphuret, and the oxygen with the metallic antimony, both of which are dissolved, so that the solution may be regarded as a compound of sulphuretted hydrogen, oxyde of antimony, and potash. On adding the diluted acid, sulphate of potash is formed, which remains in solution, and the hydrosulphuret of the oxyde of antimony is precipitated. Such is the general nature of the reaction that probably takes place, though some suppose that the precipitate thrown down by the sulphuric acid, is composed of sulphuret of antimony and water, the sulphuretted hydrogen and oxyde of antimony reacting on each other when the potash is withdrawn, and producing these compounds.

The best proportions for making the precipitated sulphurets, are four parts, by weight, of the saturated solution of potash, three of water, and two of sulphuret of antimony; or we may take equal weights of sulphuret of antimony and fused potash, and boil these in eight times their weight of water; an excess of sulphuret does no harm, as it is not dissolved, and it affords a larger surface for the action of the potash and water.

When the solution is cooled without adding an acid, a large quantity of a similar precipitate is slowly deposited, the solution of potash not being able to dissolve so much when it is cold. This precipitate is the *kermes mineral* of the shops, an article formerly employed very extensively in the practice of medicine. It appears, by analysis, to consist of sulphuretted hydrogen 17, or one equivalent, and protoxyde of antimony 52, or one equiv-

alent, and is, therefore, a *hydrosulphuret of antimony*. Formerly, it was termed *hydrosulphuretum stibii rubrum*, or the red hydrosulphuret of antimony. The name *kermes* was given, on account of the resemblance of the precipitate, in point of color, to the insect of that name.

If the liquid, after the precipitation of the kermes, is filtered, and diluted sulphuric acid added, an additional precipitate is thrown down. The potash, having retained a considerable quantity of protoxyde of antimony, over and above that which fell spontaneously, the sulphuric acid is added to destroy this union, by combining with the alkali. The protoxyde, thus liberated from its connexion with the potash, is precipitated, in combination with sulphuretted hydrogen, constituting the *sulphuretted hydrosulphuret of antimony*, or the *golden sulphuret of antimony*. It is composed of one equivalent of bi-sulphuretted hydrogen, or 33, and one equivalent of protoxyde of antimony, or 52. The actual difference between this article and the kermes mineral, seems to consist, in the larger quantity of sulphur contained in the golden sulphuret; in other respects, their composition is the same.

The golden sulphuret has been largely employed in medicine, but at present other articles are preferred. The *incompatibles* of these precipitated sulphurets are, the *acids* and *acidulous salts*, inasmuch as, by dissolving the protoxyde of antimony, they increase the emetic action, and so frustrate the specific object in view. When, therefore, acid is suspected to exist in the stomach, these antimonial precipitates should be combined with magnesia, in order to neutralise the acid.

I have observed, that the common sulphuret of antimony, or the crude antimony of the shops, was used only as a medicine for cattle. It is proper, however, to say, that cups were formerly made out of solid lumps of the sulphuret, which were kept full of wine. After a short period, the fluid was found to be so much impregnated with the metal or its protoxyde, as to have decidedly emetic properties. It is probable, that this was the first mode of employing antimony as an emetic.

As this sulphuret is employed for making several important medicinal preparations, it should be as free of impurities, as possible. *Lead* is sometimes combined with it, giving to the mass a foliated appearance, and not being convertible into vapors at a red heat, as antimony is. *Arsenic* is sometimes present, and may be detected, by the garlic odor emitted when the powder is thrown on burning coals, and from the use of the other appropriate tests.\*

\* See the chapter on arsenic.

*Sulphate of Antimony.* This salt may be made, by the action of sulphuric acid on metallic antimony, reduced to a very fine powder. The proportions are, one part of metal, and two, by weight, of acid; the process should be conducted in an iron vessel, over an open fire, and an iron rod should be employed to stir the mixture frequently. The metal takes sufficient oxygen from the acid, to be converted into a protoxyde; and the acid, thus deprived of a portion of its oxygen, is changed into sulphurous acid, which passes off in form of sulphurous acid gas. The protoxyde then combines with some undecomposed acid, and a sulphate of antimony is the result, as the diagram makes manifest.

40 or 1 equiv. sulphuric acid	{ sulphur 16 oxygen 8 oxygen 8 oxygen 8	32 sulphurous acid.
44 or 1 equiv. antimony	44	
40 or 1 equiv. sulphuric ac.	40	92 sulph. of antimony.
<hr/> 124	<hr/> 124	<hr/> 124

*Tartrate of Antimony and Potash.* This is by far the most interesting of the preparations of antimony, and merits the particular attention of every student. It is known by the names of *tartar emetic*, *stibiated tartar*, *tartarised antimony*, *tartrate of antimony*, &c. It is supposed to have been known, at so early a period as the year 1620, as some account of a preparation, in all material respects like the tartar emetic, is contained in a book published at that time, with the title of '*Methodus in Pulverum.*'

This salt is now generally regarded as a triple compound, although a binary salt, called tartrate of antimony, is mentioned in the books. The latter differs from the tartar emetic, in the difficulty with which it crystallises, and also in its solution assuming a gelatinous form.

The tartrate of antimony and potash is prepared, by boiling together equal weights of cream of tartar and oxyde of antimony, in four times their weight of water, filtering the solution afterwards, and evaporating the clear liquid till a pellicle appears on its surface. Both the cream of tartar and the oxyde of antimony should be reduced to a fine powder before they are mixed, and the water made to boil before they are put in; the boiling must be continued at least for half an hour.

The protoxyde, obtained by precipitation from the muriate, is perhaps preferable to that prepared in any of the other methods, as it requires little trouble to reduce it to a minute

state of division; the glass of antimony is not so easily reduced to an impalpable powder. The oxyde, prepared as Mr Phillips directs, from the sulphate, succeeds very well, if proper attention is paid to the pounding of the metallic antimony before digesting it in the acid, and the separation of the subsulphate from any metallic antimony that may not have been oxydated.

Though equal weights of the oxyde and cream of tartar are recommended to be used, the whole of the oxyde is not dissolved. One equivalent of cream of tartar, 198, which contains 18 parts, or two equivalents of water, combines with three equivalents of the protoxyde of antimony to form tartar emetic, and the crystals contain three equivalents of water; the chemical equivalent of the crystallised tartrate must accordingly be 363.

Tartaric acid ( $66 \times 2$ )	132
Potash	48
Protoxyde of antimony ( $52 \times 3$ )	156
Water ( $9 \times 3$ )	27
	<hr/>
	363
	<hr/>

Tartar emetic is soluble in three times its weight of boiling water, and in fifteen parts of cold water; its solution in water is decomposed by a number of acids, by alkalis and alkaline earths, and by a great many vegetable infusions and decoctions.

*Exp.* Add some nitric, sulphuric, or muriatic acid, to a strong solution of tartar emetic; a copious precipitate is thrown down, consisting principally of cream of tartar.

To a similar solution, add a small quantity of a solution of potash, soda, or ammonia. The alkali combines with the tartaric acid, and oxyde of antimony is precipitated.

A precipitate is also thrown down by lime water, and by a solution of barytes; it consists of oxyde of antimony and tartrate of lime or barytes.

*Exp.* When a solution of the hydrosulphuret of ammonia or of any other soluble hydrosulphuret, is added to a solution of tartar emetic, or of any other salt of antimony, a copious precipitate is thrown down, similar to what is obtained on adding sulphuric acid to the solution of the hydrosulphuret of antimony and potash; the same reaction taking place between the sulphuretted hydrogen and the oxyde of antimony, that has been already described.

When too large a dose of this salt has been given, the best antidotes are, sulphuretted hydrogen water, solutions of the hydrosulphurets, and infusions of cinchona bark, galls, or tea.

In a case treated by Dr Duncan, where a dose sufficient to have caused death had been taken, he gave a solution of the *hepar sulphuris* with complete success.

The most delicate test of tartar emetic, in solution, is sulphuretted hydrogen; it produces a very deep and characteristic reddish-brown precipitate, even in solutions containing a very minute quantity of this salt. The student should now transmit a stream of sulphuretted hydrogen through solutions of tartar emetic, of different strengths, and mixed with a number of other liquids, as milk, tea, porter, &c., till he becomes familiar with the appearance it presents.

In applying this test for the detection of antimony in mixed solutions, where it may have been suspected to have been administered as a poison, Dr Turner recommends the liquid to be boiled for a few minutes with a drachm or two of muriatic and tartaric acids before filtration; the tartaric acid retaining the antimony in solution, and dissolving any of the oxyde that may have been precipitated by an infusion of tea, bark, or galls, while the muriatic acid promotes the coagulation of any caseous matter that may be present. The metallic sulphuret is then to be dried and put into a small glass tube, about three inches long and a quarter of an inch in diameter, transmitting a stream of hydrogen gas slowly over it, by connecting one end of the tube by a cork with an apparatus from which this gas is disengaged, and adapting a bent tube to the open end, the other extremity of which is made to dip under water. The tube must be heated by a spirit lamp at the part where the sulphuret is placed, and great care must be taken not to apply the heat till the air of the apparatus has been completely expelled, to prevent an explosion taking place. Towards the end of the process, the temperature ought to be increased to bright redness, by directing the flame of the lamp upon the part of the tube immediately below the sulphuret, with the blowpipe.

By this process the sulphuret is completely decomposed, the hydrogen combining with the sulphur and forming sulphuretted hydrogen, while the metallic antimony is melted, and forms a film upon the internal surface of the glass, or collects in minute globules. A light green glass tube should always be preferred, containing no lead, and therefore, being less easily softened.

Throw some tartar emetic upon red hot cinders; the tartaric acid will be completely decomposed, and the carbon combining with part of the oxygen of the oxyde, metallic antimony is set at liberty, and appears in the form of minute globules. Exp.

By exposing a very small quantity to heat, on charcoal, before



the flame of the blowpipe, the metallic antimony is seen more distinctly, and white fumes of the oxyde of antimony are at the same time produced.

Exp. In addition to what we have said of the decomposition of this salt, we further notice, that the *mineral acids*, *alkalis*, and their *carbonates*, most of the *metals*, the *soaps*, and the *infusions* and *decoctions* of many *bitter* and *astringent vegetables*, are *incompatible* with the use of tartar emetic. One ounce of the decoction of yellow Peruvian bark, made in the usual way, will entirely decompose a scruple of this salt and render it entirely inert; or, if the mixture possess any power, it will be exhibited in its cathartic action. On this principle, perhaps, is based the *bolus for quartan fevers*, formerly employed by the French physicians, consisting of tartar emetic and Peruvian bark, and which usually caused an evacuation of the bowels. I know, from frequent use of the mixture, that tartar emetic and sulphate of quinine may be administered, without producing any emetic operation; the combination seems to change the ordinary action of the tartar emetic, so as to make it, strictly, an antiphlogistic, or sedative remedy, to a certain extent, and thus qualify the tonic and stimulant operation of the sulphate of quinine.

On the above principle, decoctions of bark and solutions of sulphate of quinine have been found useful in counteracting the tendencies of an overdose of tartar emetic. Infusions and tincture of galls and of rhubarb have been used for the same purpose, while gentian and worinwood do not in the least degree neutralise or destroy the poison.

As a general rule, a patient laboring under the effects of an excessive portion of tartar emetic, should swallow large draughts of sugared water, to excite free vomiting, after which, the above decoctions should be given, frequently and copiously. No irritating emetics should be employed in a case of this sort, as they always aggravate the symptoms.

The *alkaline sulphates*, as Glauber's salts, produce no change in tartar emetic, but if an excess of acid be present, constituting bi-sulphates, that excess will occasion a precipitate of insoluble sulphate of antimony. The *emeto-purgative* of the French, was composed of emetic tartar and sulphate of soda, in solution, and although regarded by some as an unchemical mixture, is not repugnant to any known principle of combination. The salts being soluble in water, are both held in solution alike, and each article produces its appropriate effects.

Tartar emetic is often adulterated. It ought to be purchased in the crystalline form, and the crystals should be white, without

odor, and having a slightly styptic and metallic taste. On exposure to the air, the crystals slightly effloresce, and if thrown on burning coals, they become black, and yield metallic antimony. A solution of the pure crystals should give a copious, gold colored precipitate with sulphuret of ammonia. The addition of acetate of lead should occasion a precipitate, soluble in nitric acid; and lime water should throw down a white and very thick mass, readily soluble in pure nitric acid. If the crystals deliquesce rapidly, in the air, the presence of other salts may be inferred.

The uses of tartar emetic, in medicine, are very extensive. Externally, it forms the basis of an active ointment, employed as a counter irritant. It forms the basis of the *wine or liquor of tartarised antimony*, each half (fluid) ounce of which contains one grain of the tartarised antimony.

Dr Paris states some very important facts, in relation to the liquor or wine of tartar emetic, (or as it is more commonly called *antimonial wine*,) which merit serious attention, because this article is one of very extensive use. He tells us, that on examination of the retail apothecary shops in London, he found, in many instances, that the protoxyde of antimony was entirely precipitated, and collected at the bottom of the vessel. So complete was the separation, in some instances, that not a trace of antimonial salts could be detected by the usual tests. Of course, the mixture (for it could not be called medicine) was positively inert, and it is easy to conjecture the evils likely to result. We are further informed, that this difficulty arose from the use of an inferior wine, as the solvent of the tartar emetic. When *sound Sherry* was employed, an efficient and permanent solution was obtained, and although a precipitation was apparent, it contained no antimonial oxyde, but simply the tartrate of lime, an incidental impurity derived from the cream of tartar, used in forming the tartarised antimony. To obviate the evils attendant on the use of bad wines, the colleges now direct a small quantity of rectified spirit to be added to distilled water, or if wine be employed, the spirit is added to that.

If the solution be made by using rectified spirit and water, it is properly called *liquor antimonii tartarisi*; but if the wine be the solvent, the appellation of *vinum antimonii tartarisi*, should be given.

*Muriate of Antimony.* This salt is frequently used, to furnish the protoxyde, by precipitation. It is made, by boiling two ounces of sulphuret of antimony, in eleven ounces of muriatic acid, in a glass vessel, for the space of an hour. The solution thus prepared, always contains a slight excess of acid, and

should be filtered, when cold, through a double linen cloth, or be permitted to remain at rest, till all the particles of undecomposed sulphuret of antimony shall have been deposited. If the oxydation of the antimony be not effected, with sufficient celerity, by the oxygen of the water contained in the muriatic acid, a drachm of nitric acid should be poured in. The following diagram exhibits the action that ensues, when common muriatic acid is employed; every equivalent of sulphuret of antimony decomposes one of water, and a corresponding quantity of sulphuretted hydrogen gas and oxyde of antimony is formed, the latter remaining in combination with the muriatic acid.

9 or 1 equiv.	{hydrogen	1	.....	17 sulphuretted hyd.
of water	{water	8	.....	
60 or 1 equiv.	{sulphur	16	.....	
of sulph. ant.	{antimony	44	.....	
37 or 1 equiv. of mur. acid		37	-----	89 muriate of ant.
<hr/> 106 <hr/>		<hr/> 106 <hr/>		<hr/> 106 <hr/>

## CHAPTER VI.

### Of Arsenic.

IN speaking of arsenic, ordinarily, a different article is meant from that now to be considered. We mean by that term, at present, the pure metallic substance, which is the radical base of all arsenical preparations.

Metallic arsenic is, occasionally, found nearly in a pure state, but more frequently in connexion with other metals, as cobalt and iron. On roasting the ores of arsenic in a reverberatory furnace, the arsenic, by reason of its volatile character, is driven off, combining with the oxygen of the air as it rises, and condensing in thick cakes, on the sides of the chimney. The sublimed mass, is resublimed, and the product is the virulent poison, known by the names of *arsenic*, *white oxyde of arsenic*, *arsenious acid*, and *ratsbane*. From this substance, the metal may be detached, by the deoxydating power of charcoal.

Its equivalent is 38, and it is volatilised at 390°. Some

diversity of opinion obtains, in relation to its specific gravity, but it is, perhaps, most correctly stated at 5.7.

The most convenient method for reducing the metal, is to expose the white oxyde to heat, in mixture with finely powdered charcoal. The oxygen of the oxyde unites with the carbon, and forms carbonic acid, which flies off in the state of gas, while the metal is at the same time volatilised, and should be condensed in a close vessel, so as to prevent the action of the external air. When a small quantity is required, for experimental use, the best mode will be, to mix the oxyde, intimately, with twice its weight of black flux, having them both previously dry. The mixture should be exposed to heat in a crucible, having another luted over it, to collect the product, leaving a small aperture for the escape of gas. The lower crucible should be placed in a sand bath furnace, and the upper one kept as cool as possible, and completely out of the sand, that the arsenic may condense; the process may be easily conducted with a small chauffer or furnace, taking care always to protect the upper crucible from the heat, as much as possible.

Instead of using a crucible, the reduction may be effected more easily, in operating with small quantities, by heating the mixture in a glass tube, held in the flame of a spirit lamp. The tube should be perfectly dry, and the mixture placed in a small piece of paper, so as to slide readily down the tube, when held in an inclined position. The student should repeat this process frequently, in tubes of different sizes, from a half to a quarter of an inch in diameter, and from two to four inches long, using different quantities of the mixture, from one-eighth of a grain to seven or eight grains, till he has become quite familiar with the appearance the metallic arsenic presents.

The black flux is preferred to pure charcoal, as its potash is supposed to prevent the arsenious acid from being volatilised, retaining it till it is decomposed by the charcoal which is mixed with it. When the crust is to be removed from the tube, a file should be drawn across it, immediately below the part where the metallic arsenic has collected; a slight pressure will then be sufficient to break off the lower part of the tube, and the metallic arsenic may be separated by a penknife. It has a steel-gray color, and is very brittle. It exhibits a metallic lustre, and a crystalline structure.

Some of the books tell us, that the white oxyde of arsenic (arsenious acid) gives the smell of garlic, when thrown on red hot coals; but if that be the case at all, it must depend on the previous reduction of the metal from the oxyde, by the agency of the carbonaceous matter with which it is in contact. The



oxyde, as such, cannot give that indication under any circumstances.

There are two compounds resulting from the union of oxygen with arsenic. These have been called the deutoxyde and peroxyde; but as they are both possessed of acid properties, we shall consider them as acids, and not as oxydes. They are the *arsenious* and *arsenic acids*, and their salts are termed *arsenites* and *arseniates*.

The composition of these acids is stated by Reid, as follows.

	Arsenic.	Oxygen.
Arsenious acid	38 or 1 equivalent	+ 16 = 54
Arsenic acid	38 or 1 equivalent	+ 24 = 62

Turner makes the proportion of oxygen, in the first, to be 12, and in the last, 20.

We notice *arsenious acid* first, because vastly more important than the arsenic acid, in every point of view. It is always formed, when the metallic arsenic is heated in open vessels, or oxydated by means of diluted nitric acid. At a heat below redness, it is volatilised, but the vapors have not the alliaceous smell peculiar to the pure metal, and when condensed on cold surfaces, the acid is recovered.

When arsenious acid is suddenly heated beyond the point at which it sublimes, it fuses into a transparent, brittle glass, which becomes opaque by keeping. The specific gravity of this glass is about 3.7.

A good deal of diversity has obtained, concerning the taste of arsenious acid; some persons affirming that it is quite acrid, others that it excites a faint impression of sweetness, and others that it has an acid taste. It may be, that this variety has been occasioned by the adventitious presence of some unknown agent. The samples I have tasted, communicated no very distinct impression of any sort, nor do I believe, that the pure article has any property very obvious to the taste. It reddens vegetable blues rather feebly; in order to produce this effect, litmus paper should be moistened.

Arsenious acid, or the white oxyde of arsenic, has long been employed in medical practice, both internally and externally. Internally, it is used in form of pill, and also in the shape of what is called the *liquor arsenicalis* or the *mineral solution of Fowler*. Externally, it has been applied by all sorts of cancer quacks, to cure cancerous, scrophulous, and other sores. The celebrated *pate arsenicale* of the French, owed its activity to the arsenious acid.

The *mineral solution* or the *liquor arsenicalis* is prepared, by boiling 64 grains of carbonate of potash, and as much of the



white oxyde, in six or eight ounces of water, until the solid matter is all taken up. When cold, an ounce of compound spirit of lavender is to be added, and as much water as will make the whole equal to one pint. The solution contains arsenite of potash.

The origin of the use of arsenic in fevers, is ascribed, by some, to the preservative power displayed by its vapors in the vicinity of copper works, where these vapors are evolved. Thus it is affirmed, that previous to the establishment of the Cornwall copper works, the marshes in the neighborhood were constantly producing intermittent fever, but since that epoch, a case of that disease has not occurred; and it is a common remark with the workmen, that the smoke *kills* all fevers. It is also stated, that no agricultural or other improvements in the vicinity, can account for this important change.

According to Orfila, the long continued internal use of the solution of arsenic, as well as the external application of the various ointments containing arsenious acid, has occasioned all the symptoms of poisoning by arsenic. All the writers agree, that the application of arsenic to an abraded surface, is more rapidly fatal than the internal administration of this poison. For some interesting details, on this subject, the reader may consult Christison, Paris, and other authorities.

Arsenious acid is often administered as a poison, and it has been sold and taken, in mistake. In appearance, it resembles magnesia, and has actually been swallowed for that article. The necessity of caution in apothecary shops and families, in regard to this substance, cannot be too strongly enforced; it should always be marked, in perfectly legible characters, *poison*.

Diversity of opinion has prevailed, respecting the length of time after death, in which this poison may be satisfactorily detected. In the case of Mina, in Pennsylvania, the death of his victim, by arsenic, was clearly proven, although interment had taken place three months before. In the seventeenth volume of the *Lancet*, the reader may find a very interesting case, of the detection of arsenic in the body of an individual, who had been in the grave seven years.

These facts show the indestructibility of arsenic, though directly exposed to extensive decompositions of matter contiguous to it. By detaching the stomach and intestines from the body, and examining their contents, by the methods presently to be stated, we can hardly fail to detect arsenic, if it be really present. But if an individual be destroyed by the poisonous action of arsenic, externally applied, we are left entirely to the symptoms and history of the case, to form our judgment, for analysis cannot aid us.

As arsenious acid is the most important compound of arsenic, and as it is frequently administered as a poison, the student should perform a number of experiments, to render himself familiar with the appearances which the different reagents, that have been recommended for detecting it, present with arsenious acid, when pure, or mixed with a number of other substances, such as are usually found in the contents of the stomach, or are likely to have been given at the same time, either as an antidote, or for any other purpose.

Make a solution of arsenious acid, reducing it to fine powder, and boiling it for twenty minutes, in twenty-five or thirty parts of water. Sixteen parts of boiling water are said to dissolve one of arsenious acid, five-eighths of which are deposited in the form of crystals as it cools, the rest remaining in solution. When arsenious acid is merely mixed with water, at natural temperatures, the solution does not contain more than a four-hundredth part of arsenious acid.

Mix equal weights of arsenious acid and carbonate of potash, and boil the mixture in ten or twelve times its weight of water; the arsenious acid unites with part of the potash, and is very speedily dissolved, the arsenite of potash being much more soluble than pure arsenious acid; as the excess of alkali which the solution contains, (when the materials have been mixed in the proper proportion,) must react upon many of the tests used for the detection of arsenious acid, the student must always bear in mind, the precise composition of the fluid on which he is operating.

Mix a few drops of the saturated solution of arsenious acid, with seven or eight ounces of water, and pass a stream of sulphuretted hydrogen through the liquid. The sulphur of the sulphuretted hydrogen, combines with the metallic arsenic, rendering the liquid of a yellow color, but very faint, from the small quantity of materials present, and the hydrogen combines with the oxygen of the arsenious acid, forming water. Boil the liquid to expel the excess of sulphuretted hydrogen, and then set it aside for several hours, when a minute quantity of sulphuret of arsenic will be gradually deposited at the bottom. The sulphuretted hydrogen may be prepared in a bottle with a bent tube fitted to it, the other extremity being introduced into a glass containing the solution.

Pass a stream of sulphuretted hydrogen gas through a strong solution of arsenious acid, in water; observe the large quantity of the sulphuret of arsenic which is precipitated, of a rich yellow color; place it on a filter, when no farther precipitation takes place, and after washing it several times with water, set it aside, that it may dry.

Transmit a stream of sulphuretted hydrogen through a solution of the arsenite of potash, taking care to have a slight excess of alkali. No precipitate is thrown down, nor does the gas appear to have any effect upon the liquid; sulphuretted hydrogen gas not being capable of decomposing arsenious acid, when in combination with potash. Add an excess of muriatic acid, and immediately, the characteristic, yellow colored precipitate will appear. Exp.

Put a few drops of the solution of arsenious acid in water, into an ounce or two of water, and add a small quantity of a solution of the sulphate, nitrate, or acetate of copper. The liquid remains quite transparent and colorless, the arsenious acid not having so great an affinity for the oxyde of copper, as the acid with which it is already combined. If a small quantity of an alkaline solution be now added, the alkali will unite with the acid of the salt employed, and remain in solution, and the arsenious acid, combining with the oxyde of copper, will form arsenite of copper, which is insoluble in water, and is precipitated of a grass-green color. If the arsenious acid shall have been previously combined with potash, the grass-green precipitate appears immediately. Exp.

Mix a drop or two of the solution of arsenite of potash, with a solution of the sulphate of copper in a glass of water. If a great excess of potash be employed, the oxyde of copper alone will be precipitated, and of a blue color, very different from the grass-green precipitate of arsenite of copper. Add a solution of potash to a solution of the sulphate of copper, and compare the color of the precipitate, with that of precipitated arsenite of copper. Exp.

Instead of using potash to combine with the acid of the salt of copper, and allowing the arsenious acid to unite with the oxyde, a solution of ammonia is frequently employed, and indeed it is to be preferred, as it is easy to combine the salt of copper with the exact quantity of ammonia that may be required for the precipitation of the arsenious acid; the precipitate that is thrown down in this case, however, has not such a rich, green color, as when potash is used. Exp.

To prepare a solution of a salt of copper for this purpose, ammonia must be added to it, till the precipitate, that is thrown down at first, is almost entirely redissolved, decanting the deep blue colored liquid that is obtained in this manner, and keeping it in a bottle accurately closed; otherwise, the ammonia will soon escape, on exposure to the air. The ammonia unites with the acid of the salt, forming another salt, which remains in solution; and the precipitate that is thrown Exp.

down, consists of oxyde of copper, which is redissolved by the ammonia that is afterwards added. If more ammonia be used than is sufficient to redissolve the whole of the precipitated oxyde, the solution will not give any precipitate with the arsenious acid, the arsenite of copper being soluble in an excess of ammonia.

Add a few drops of a solution of the ammoniaco-nitrate of copper, prepared in the manner described, to a solution of arsenious acid. Arsenite of copper will be immediately precipitated. Diffuse the precipitate through the liquid, and divide it into two portions; then add a little nitric acid to the one and some ammonia to the other; the precipitate in each will be redissolved, the arsenite of copper being soluble both in nitric acid and ammonia.

Drop a solution of nitrate of silver into a solution of arsenious acid in distilled water. No precipitate is thrown down, nitric acid having a stronger affinity for oxyde of silver, than arsenious acid; if a little potash be now added, it combines with the nitric acid and forms nitrate of potash, which remains in solution, and the arsenious acid, combining with the oxyde, forms a yellow colored precipitate, which is arsenite of silver.

Add a solution of the nitrate of silver to a solution of phosphate of soda. Phosphate of silver is precipitated, of a yellow color, and nitrate of soda remains in solution. The nitrate of silver cannot therefore be used as a test of the presence of arsenious acid in solutions which may be suspected to contain phosphate of soda, as in liquids obtained from the stomachs of persons supposed to have been poisoned by arsenic.

Prepare a solution of the ammoniaco-nitrate of silver, by adding ammonia in small quantities at a time, to a solution of the nitrate of silver, proceeding in the same manner as in the preparation of ammoniaco-nitrate of copper. Then drop a little into a very diluted solution of arsenious acid; the ammonia remains in combination with the nitric acid, and the arsenious acid, combining with the oxyde, gives the characteristic, yellow colored precipitate of arsenite of silver.

If the ammoniaco-nitrate be mixed with a solution of the phosphate of soda, a white precipitate will be thrown down, instead of the yellow-colored precipitate, which the nitrate of silver gives with a solution of this salt; and accordingly, the ammoniaco-nitrate of silver is always preferred to the nitrate, in testing any liquid for the presence of this poison.

Precipitate some arsenite of silver from a solution of arsen-

ious acid by the ammoniaco-nitrate of silver, diffuse the precipitate through the liquid, divide it into two portions, and add ammonia to one and nitric acid to the other; both will be redissolved, and accordingly, great care must be taken to have no excess either of acid or alkali, in using the nitrate of silver, as a test of the presence of arsenious acid, otherwise no precipitate will appear, even though a considerable quantity of arsenious acid should exist in solution. Exp.

Mix some lime water with a small quantity of a solution of arsenious acid; arsenite of lime is immediately precipitated, in the form of a white powder. Exp.

Put a few drops of a solution of the bi-chromate of potash\* into a solution of arsenious acid. The liquid will assume a rich, pea-green color, after standing for some time; heat a little of it by a spirit lamp, and the green color will be developed immediately. The change of color is owing to the arsenious acid attracting oxygen from part of the chromic acid, and converting it into oxyde of chrome. Exp.

Drop a little of the solution of bi-chromate of potash into a solution of tartar emetic; the liquid will assume the same green color as in the preceding experiment, a circumstance that was pointed out by Mr Lawrence Reid. It follows, therefore, that the bi-chromate of potash cannot be used as a test of arsenic, in any solution, which may be suspected to contain tartar emetic. Exp.

The student, having now made himself familiar with the appearances which the most important tests for the detection of arsenious acid produce, when mixed with a solution of this substance, and the precautions which he must take in applying them, should perform a number of experiments, in the next place, with liquids containing animal and vegetable matter, and mixed with arsenious acid, both in solution and in the solid form, till he is able to detect it, when the liquid upon which he is operating, contains only a very minute portion of arsenious acid. He must recollect, however, that in operating with mixed liquids, such as are generally met with in cases of poisoning by arsenious acid, where a variety of animal and vegetable principles are generally blended together, and where various kinds of saline matter are also likely to be present, he cannot expect that he will be able to recognise it, so easily, as in a solution of pure arsenious acid in water. It is now admitted, indeed, that we cannot depend upon the appearances which

\* This test was introduced by Professor Cooper, of the University of South Carolina.



any of the tests present, when mixed with these liquids, as unequivocal indications of the presence of arsenious acid; as it has been proved, that they often fail in producing the characteristic precipitates in such compound fluids, though arsenious acid may be present, and occasionally, they cause the same appearance as when arsenious acid is present, though the liquid does not contain any. A number of important experiments and observations on this point are stated in the *Edinburgh Medical and Surgical Journal*, for July, 1824, in an excellent paper by Dr Christison.

Though considerable information may be obtained, by applying the tests we have described to mixed solutions, suspected to contain the poison, and when they all concur, in the indications which they give of arsenious acid, little doubt can be entertained of its presence; still, in order to avoid every source of fallacy, it will be necessary to continue our investigation still farther, separating the matter that appears to have produced the characteristic precipitate with the arsenious acid, and extracting the metal itself, if any arsenious acid shall have been present.

For this purpose, Dr Christison recommends sulphuretted hydrogen to be employed, dispensing with the other tests; the liquid should be boiled and filtered in the first place, and then acidulated with muriatic or acetic acid, to prevent any alkaline matter, that may be present, from interfering with the precipitation; after which, a stream of sulphuretted hydrogen should be passed through the liquid, and continued, at least, for half an hour; it is then to be boiled, for a few minutes, to expel any excess of sulphuretted hydrogen, and the precipitate collected on a filter, washing it repeatedly with water, and drying it afterwards, by heating it at a temperature of  $212^{\circ}$ . On mixing it intimately with about twice its weight of black flux, and exposing it to heat in a glass tube, over a spirit lamp, the potassium in the black flux combines with the sulphur, and the metallic arsenic is sublimed in the same manner, as in the reduction of arsenious acid, by the same substance. The size of the tube must be adapted to the quantity of the precipitate which has been procured; the most convenient size is about three or four inches long, and about a quarter or a half inch in diameter; the mixture should not fill more than half an inch of the lower part of the tube, and smaller tubes should be used, when only a very minute quantity of matter has been precipitated.

If a crust of metallic arsenic should be obtained, its steel-gray lustre, its brittleness, the facility with which it is volatilised, and the garlic odor that is, at the same time, produced, will be suffi-

cient to distinguish it from any other substance; if, however, there are only very indistinct appearances of the metallic arsenic, the following is the method that I have found most satisfactory, for ascertaining if arsenic be really present. The tube is to be exposed again to heat, over the spirit lamp, till the matter that has been sublimed, is carried a little farther up the tube and completely separated from the black matter that remains at the bottom; the lower part of the tube must then be broken off, drawing a file across it, previously, that it may be easily removed, and the upper part put into another glass tube and boiled, for five or ten minutes, with a little water, to which a few drops of nitric acid have been added. If the tube Exp. be coated with any arsenious acid, it will be immediately dissolved, and if any metallic arsenic should be present, it also will be converted into arsenious acid, attracting oxygen from the nitric acid, and being dissolved at the same time, so that the liquid may now be considered as a solution of arsenious acid in water, with a small quantity of nitric acid; and accordingly, on neutralising the excess of acid, by dropping ammonia into it, through a test-tube drawn out at the extremity over a spirit lamp, so as to represent a small funnel terminating in a capillary tube, the ammoniaco-nitrates of copper and silver will produce the characteristic green and yellow precipitates. If they should give no precipitates, then we may conclude, that the mixture which we have examined, contains no arsenic.

Digest a quarter of a grain of metallic arsenic in a test-tube with a drachm of water and two or three drops of nitric acid, and neutralise the solution with ammonia, in the manner described in the preceding paragraph. Put a number of drops of the solution, on different parts of a sheet of white paper, and touch them with glass rods, dipped in the different solutions used for the detection of arsenious acid, and the characteristic color will appear in each. We must be careful not to dip the same glass rods into different solutions, without previously removing any liquid that may be still adhering to them. Exp.

The indications, given by the different tests, are free from every source of fallacy, when they are added to solutions prepared in this manner; as the arsenious acid is thus removed from all the animal and vegetable matter, that might have interfered with their action in the mixed liquid. Infusions of astringent matter and some other vegetable and animal substances, which are frequently met with in the liquid contents of the stomach, have been shown, by Dr Christison, to be capable of retaining, in solution, the precipitates that arsenious acid gives with the

ammoniaco-nitrates of copper and silver. Tartaric and acetic acids appear to be solvents, in some cases, arsenite of silver being soluble in both.

Orfila has proposed to use chlorine, and Mr Phillips animal charcoal, to decolorise mixed fluids, suspected to contain arsenious acid, and allow the usual tests to be applied in the liquid way, without previously removing the arsenious acid and subjecting it to some process of reduction; but though both may occasionally be used with advantage, it will be better to adopt the method we have already mentioned. It appears, also, from the experiments of Drs Christison and Paris, that charcoal is capable of precipitating arsenious acid from its solution in water.

If a quantity of solid powder be obtained among the contents of the stomach, the following experiments may be made.

Expose a small quantity to heat, on a thin plate of copper, or on the blade of a knife, over a spirit lamp; it will be completely volatilised if it be arsenious acid. Mix another portion with twice its weight of black flux, and if any crust be obtained which

Exp. has the character of metallic arsenic, it will be unnecessary to proceed any farther. Should the appearance of the crust be unsatisfactory, it must be treated in the manner described in page 437, and if no indication of arsenious acid be then obtained, we may conclude that the white powder does not contain any.

I have frequently examined portions of matter, supposed to contain arsenic; and although the usual tests have been so satisfactory as to leave no doubt in my own mind, I have never trusted to them alone, but have always considered myself bound to resort to the process for metallic reduction. If we first obtain the metallic crust, and then, in the manner described, prove this crust to be arsenical, we have no reason to doubt, for a single moment. In all our examinations, however, we are to bear in mind, that the life of a fellow creature may be suspended on our decision.

Some diversity of opinion exists, in regard to the quantity of arsenic that will prove fatal. But it is not possible to settle this question by any uniform rule, since a variety of contingencies may concur, to give a difference of result, under circumstances apparently the same. Hence some authors tell us, that two grains of the white oxyde have destroyed life, while others assert, that larger doses have done, comparatively, no injury.

When arsenic is given as a poison, the dose is not limited by any rule, other than the rule of extermination, and in all such cases, large portions are swallowed. Occasionally, bad con-

sequences have followed the practice of attempting to kill rats with arsenic, and though, fortunately, the poison is not then received in quantities sufficiently large to be fatal to those who, unwittingly, are injured by this article, we think the custom should be entirely discarded. Happily we are in possession of a very good substitute; which, while it destroys the vermin, is not likely to be deleterious to man or other animals. If plaster of Paris (sulphate of lime) be reduced to fine powder, and well mixed with an equal quantity of oat-meal or the flour of Indian corn, and placed in the cellar or other place, infested with rats, they will soon disappear. The mixture is greedily eaten, and its dryness excites the animal to drink water. As soon as the plaster is moistened, it *sets* or becomes compact, and finally grows so hard that it cannot be evacuated. It is affirmed, that the rat is the only animal that is known to eat this mixture.\*

It would be well, if every state government were to prohibit the sale of arsenic in any form whatever, excepting to physicians, or to individuals furnished with regular medical prescriptions; and assuredly, if the article, above stated, be capable of destroying rats, it should be made a criminal offence, to sell any quantity of arsenic under the mere pretext of administering *ratsbane*.

*Arsenic acid* is prepared by digesting metallic arsenic, or arsenious acid in strong nitric acid, mixed with a little muriatic acid, evaporating the solution afterwards to dryness in a glass or earthen vessel.

If equal parts of nitre and arsenious acid are fused in a crucible, the nitric acid is completely decomposed, part of the arsenious acid attracting oxygen from it and being converted into arsenic acid, which remains in combination with the potash, while the rest is volatilised. The arseniate of potash is speedily dissolved, on digesting it in water; it gives a brick-red precipitate, with a solution of the nitrate of silver. Mix half a grain of arsenious acid with an equal weight of nitre and fuse it slowly in a platina spoon, over a spirit lamp; then put the spoon into a test-tube, containing a little water, and add a drop of a solution of the nitrate of silver to the solution of arseniate of potash thus obtained, when the characteristic, brick-colored precipitate will be thrown down.

Arsenic acid has a sour, metallic taste, reddens vegetable blues, and forms, in union with alkalis, the neutral salts that are termed *arseniates*. It is far more soluble in water than arsenious acid, dissolving in five or six times its weight of cold,

\* See London Mechanics' Magazine, for November, 1831.

and in a less quantity of hot water. When evaporated, it forms irregular grains, but does not crystallise. When strongly heated, it fuses into a deliquescent, glacial mass. Exposed to a full red heat, it is decomposed into oxygen gas and arsenious acid. It is a very active poison, but it is not resorted to as such, because it is only made for laboratory use.

Arsenic acid is decomposed by sulphuretted hydrogen gas, and yields a sulphuret of arsenic. The soluble arseniates, when mixed with nitrate of lead or silver, form arseniates that are insoluble in water, the former of which is white, and the latter, a brick-red. Both are soluble in dilute nitric acid, and if heated with charcoal, they yield metallic arsenic.

*Arseniuretted hydrogen* is a gaseous compound of arsenic and hydrogen gas, which may be prepared, by digesting an alloy of tin and arsenic, in common, liquid muriatic acid; a portion of water is decomposed, the tin uniting with its oxygen and forming oxyde of tin which remains in solution, combined with the muriatic acid, while the hydrogen unites with the arsenic, and escapes in a gaseous form. It is extremely deleterious, and has already proved fatal to M. Gehlen, a German chemist; it will be better, therefore, for the beginner to pass over this process.

Another compound of arsenic and hydrogen has been discovered by Davy; it exists in the solid form, but its exact composition has not been ascertained.

The *protosulphuret of arsenic*, or *realgar*, may be obtained by mixing arsenious acid with two-fifths of its weight of sulphur, and exposing the mixture to heat, in a covered crucible, till it is fused; part of the sulphur unites with the oxygen and the rest with the metallic arsenic, forming the deep, red colored mass that remains. It is sublimed when exposed to heat in a retort, and condenses in the neck, in the form of a very rich, red colored powder.

The *yellow sulphuret of arsenic*, or *orpiment*, contains more sulphur than the red, and may be prepared in the same manner, from a mixture of equal weights of sulphur and arsenious acid, or by transmitting sulphuretted hydrogen through a solution of this substance.

Another sulphuret of arsenic has been described, which contains more sulphur than either of the preceding compounds; it is obtained by transmitting a stream of sulphuretted hydrogen gas through a solution of arsenic acid.

The composition of the sulphurets of arsenic, is thus stated by Reid.



Sulphur. Arsenic.

Sulphuret	$16 + 38 = 54$
Sesqui-sulphuret	$24 + 38 = 62$
Bi-sulphuret	$32 + 38 = 70$

Realgar and orpiment are natural products, and are used as pigments, and also in the manufacture of shot.

The native sulphurets are less poisonous than arsenious acid; but the sulphuret formed by precipitation is more deleterious than the native sulphuret, probably because a given quantity of the latter contains so much foreign matter as to reduce or neutralise, in some degree, its power, while the precipitated article contains nothing but its essential ingredients.

*Chloride of Arsenic.* If some metallic arsenic, in fine powder, be thrown into a bottle of chlorine gas, it will take fire immediately, and a chloride of arsenic will result. A similar compound may be made, by distilling a mixture of six parts of corrosive sublimate with one of arsenic. In this case, one portion of chlorine quits the corrosive sublimate and joins the arsenic, forming chloride of arsenic, while chloride of mercury or calomel remains.

This chloride is composed of 36, or one equivalent of chlorine, and 38, or one equivalent of arsenic. It is a colorless, volatile liquid, emitting fumes when exposed to the air, and hence called *fuming liquor of arsenic*. The addition of water changes it into muriatic and arsenious acids.

The treatment of a case of poisoning by arsenic, might be inferred, to some extent, from what has been said of the modes by which it is detected. We propose, however, to pay a little further attention to this topic, as it is one of the utmost importance.

Our object will be, chiefly, to show the inefficacy of all the antidotes that have been employed, in the treatment of this poison, or rather to make manifest, that the aid expected from chemistry is, for the most part, fallacious.

It is admitted, as a general principle, that arsenic, in all its forms of combination, acts nearly in the ratio of solubility. It might be inferred, thence, pretty conclusively, that every supposed chemical remedy would be useless, that failed to render arsenic insoluble, both in water, and in the contents of the stomach and alimentary canal; but no such antidote is known. The arsenites, although insoluble in water, are so far soluble in the fluids of the stomach, as to allow the introduction of a sufficient quantity of the poison into the blood, to prove fatal. Hence the entire inutility, or rather the injurious tendency of vinegar, sugar, butter, and other oleaginous matters, lime water,

bitter decoctions, and many other articles, that have long since been discarded.\*

Nor are the sulphurets of alkalis calculated to operate as antidotes; and, in like manner, the water of sulphuretted hydrogen gas cannot counteract the poison. In both cases, a sulphuret of arsenic is produced, which although supposed to be less poisonous than arsenious acid, is sufficiently virulent to destroy life.

*Magnesia* has been extolled by Mr Hume, (an enterprising London chemist,) by Mr Buchanan, and some others, for its powers as an antidote. But as this simple and almost inert substance does not act chemically on the oxyde of arsenic, out of the body, we are not prepared to suppose, that it can decompose that substance, when it meets with it in the stomach. The remedy appears to have been given along with milk, and the recovery may have depended very much on the action of that fluid, in sheathing the mucous coat of the stomach. There can be no doubt, however, that success attended the use of *magnesia*, in the cases referred to.

*Charcoal* has also been proposed, as an antidote, and in the hands of Mr Bertrand, it has actually succeeded.† He tried the remedy, first on some inferior animals, and then on himself. He took five grains of the white oxyde at once, suspended in a charcoal emulsion, and realised no serious consequences. Now, as charcoal has the power of reducing arsenious acid to the metallic state, (in which it is inert,) we can suppose it barely possible, that such a change might take place in the human system. But even admitting this change, the reduced metal would be soon converted into an oxyde, (or rather into an acid,) unless active cathartics were employed to convey it, at once, out of the system.‡

Inasmuch as no dependence can be placed on any article that has been supposed to act on chemical principles, the great object should be, to evacuate the stomach of its contents, as speedily as possible; and hence, if vomiting be not commenced, it should be excited, at once, by the most prompt emetics, as sulphate of zinc, &c. Large quantities of mucilaginous drinks should be given at the same time, to involve and suspend the poison, until it may be dislodged.

\* Christison on Poisons.

† Journal Generale de Medicine, 1815.

‡ If, as Christison asserts, charcoal precipitates this poison from its solution in water, that may assist in the explanation of its remedial powers.

## CHAPTER VII.

## Of Tin.

THIS metal is found in Cornwall, Bohemia, Spain, and other places. It is met with in a metallic state, united with sulphur and copper; also, in the state of oxydation, along with silex and oxyde of iron, in which state it is called *tin stone*. It is from this, that the best block and grain tin of commerce, are obtained. The finest grain tin is almost chemically pure, containing very minute portions of copper and iron, and sometimes a little arsenic.

If 1200 grains of native oxyde of tin be mixed with 100 grains of finely powdered charcoal, and exposed to heat in a crucible, pure metallic tin will be obtained. Exp. The carbon combines with the oxygen of the oxyde, and escapes in the form of carbonic acid gas.

The color of tin is white, and its lustre not unlike that of silver. Exposure to the air lessens its brilliancy, but does not oxydise it, even though much moisture be present. The article called *tin-foil*, is a proof of its malleability, for it does not exceed a one-thousandth part of an inch in thickness. In point of ductility and tenacity, it ranks below other metals. It is soft and inelastic, but capable of being bent considerably without fracture. In flexing it backwards and forwards in the hands, a peculiar crackling noise is perceived,\* as though the piece were about to be broken. The specific gravity of tin is 7.5, its melting point  $442^{\circ}$ , and its equivalent 58.

The filings or powder of tin, formerly used as a remedy for worms, may be made by agitating melted tin in a heated mortar, with a heated pestle, until it cools; or the melted metal may be shaken in a wooden or iron box, well rubbed over with chalk on the inside.

The anthelmintic powers of tin filings have always been disputed, and if they have done service, it has resulted from mere mechanical action. Dr Murray regarded the article as decidedly injurious, because, as he supposed, it generated hydrogen gas in the intestines.

*Oxydes of Tin.* Tin is susceptible of two degrees of oxydation. Both of its oxydes form salts with acids, and they also combine with alkalis. The composition of the oxydes is as follows.

\* Formerly termed the *squeak* or *cry* of tin.

	Tin.	Oxygen.
Protoxyde	58, or 1 equiv.	+ 8, or 1 equiv. = 66
Peroxyde	58, “ “	+ 16, 2 equiv. = 74

Melt some tin in an iron ladle, and expose it freely to the air;   
 Exp. a crust soon gathers on its surface, composed, principally, of the *protoxyde* of tin. It may also be obtained, by adding a solution of potash to a solution of newly prepared protomuriate of tin, taking care to avoid an excess of alkali, the protoxyde being soluble in a solution of potash.

Take another portion of tin and expose it to a full white heat, in a crucible placed in a furnace, and resting about an inch and a half above the branders. It soon takes fire,   
 Exp. burning with a white flame, and combining with a larger portion of oxygen, so that it is converted into *peroxyde* of tin.

Put half an ounce of nitric acid into a deep glass, add a drachm of water to it, and pour in two or three hundred grains of the powder of tin; the metal attracts oxygen both from the acid and the water, and hydrogen and nitrogen gases, meeting in a nascent state, unite and form a portion of ammonia; this combines with a portion of nitric acid which is not decomposed, forming nitrate of ammonia, which explains the appearance of the white fumes that are mixed with the large quantity of ruddy vapors of nitrous acid, that are at the same time disengaged. Considerable heat is also produced, and the tin is left in the form of a bulky white powder, which is the *peroxyde* of tin.\*

Take a common tin plate, (which is, in reality, a plate of iron coated with tin,) about a foot square, hold it over a chauffer or before the fire, with a pair of pincers, till a drop of water, allowed to fall upon its surface, begins to boil immediately, and then wash one of its sides with a mixture of four parts, by measure, of water, one of nitric, and one of muriatic acid. Its surface will immediately assume a beautiful crystalline appearance, and by heating the plate at particular parts with the blowpipe, or exposing different parts to higher and lower temperatures, a great variety of figures may be produced, which will be seen better, on washing it with water. The *crystallised tin plate*, as it is sometimes termed, of which a great number of hardware articles are now made, is prepared in this manner; and the various colors it is made to assume, are communicated by giving it a thin coating of different colored varnishes. The

\* The compounds, formed by the oxydes of tin and metallic bases, were termed, by Berzelius, *stannates*.

tin plates used for this purpose, should have a good coating of metallic tin, otherwise most of the tin will be removed during the preparation, and nothing will remain but the iron below, which has a very dark color.

*Chlorides of Tin.* These are two in number, viz. the *chloride* and *perchloride*, and have the following composition.

	Chlorine.		Tin.	
Chloride	36, or 1 equiv.	+	58, or 1 equiv.	= 94
Perchloride	72, or 2 equiv.	+	58, " " "	= 130

The first, called also *protochloride*, may be made by evaporating the muriate of the protoxyde to dryness, and fusing the residue in a close vessel, or by heating an amalgam of tin and calomel, (chloride of mercury). It is a gray, solid substance, of a resinous lustre, fusing at a heat below redness, and changing into *per* or *bi-chloride*, when heated in chlorine gas.

*Perchloride* of tin (called also *bi-chloride*) may be prepared, by heating metallic tin or the chloride, in an atmosphere of chlorine gas; or by distilling a mixture of 8 parts of powdered tin, with 24 of corrosive sublimate. The two equivalents of chlorine in the latter, join the tin, and thus form the *per* or *bi-chloride*, while the mercury is reduced to its metallic state.

This perchloride is a colorless, volatile liquid, giving out white fumes when exposed to the atmosphere. It has a strong affinity for water, and is changed by that fluid into the permuriate. It was formerly known as the *fuming liquor of Libavius*.

*Sulphurets of Tin.* Two compounds of tin and sulphur are generally acknowledged. These are, the *protosulphuret*, and the *bi* or *persulphuret*; to these Berzelius adds another, which he calls the *sesqui-sulphuret*. The *protosulphuret* (or simply, the *sulphuret*) is composed of 16 sulphur, or one equivalent, and 58 tin, or one equivalent, whereas in the *persulphuret*, there are two equivalents of sulphur, 32, to one of tin.

The *protosulphuret* may be formed, by heating sulphur with metallic tin. The compound has a bluish-gray color and a metallic lustre; it is fusible, at a red heat, and exhibits a lamellated structure, on cooling. Muriatic acid dissolves it, with evolution of hydrogen gas.

The *persulphuret* is the compound formerly called *Mosaic gold*, or the article of which the *golden calf* was made. It is formed by heating a mixture of sulphur and peroxyde of tin, in close vessels. Formerly, the Mosaic gold was supposed to be a hydrosulphuret of the peroxyde of tin; but Dr Davy showed, most conclusively, that it is, simply, a compound of sulphur and tin, in the proportions above stated.



Sulphuric and nitric acids, are capable of dissolving tin, but the most important of its salts, is the *muriate* or *hydrochlorate* of tin.

The *muriate of tin* may be prepared by mixing tin with liquid muriatic acid, and exposing it to heat in a common oil flask; part of the tin takes oxygen from a portion of water which is decomposed, and forms oxyde of tin, which remains in combination with the muriatic acid; while another portion of the metal combines with the hydrogen, forming the fetid gas that is disengaged; it may be collected in jars over a pneumatic trough, by heating the mixture in a retort, and it burns with a blue flame. By exposing the solution of the muriate to the open air, it soon attracts oxygen, and permuriate of tin is obtained in solution. A similar solution may be procured, more speedily, by adding tin in small quantities at a time, to two parts of nitric acid, mixed with one of muriatic acid.

The solution obtained in the first instance, is the *protomuriate*, which contains 37 parts of muriatic acid, or one equivalent, and 66 parts of protoxyde of tin, also one equivalent. The permuriate is composed of 37 muriatic acid, and 74 peroxyde of tin. The protomuriate is extensively employed in the art of dyeing, particularly as a mordant. It is also used, occasionally, in the laboratory.

Tin, in solution, is detected easily, if in the form of protoxyde, by the deep purple precipitate (called the purple of Cassius) occasioned by adding a small quantity of the chloride of gold. The chloride of platina gives an orange colored precipitate; the bi-chloride of mercury a black precipitate; and the alkaline hydrosulphurets strike a dark brown color. If the solution contain peroxyde of tin, the bi-chloride of mercury will give a white precipitate, and the alkaline hydrosulphurets a golden colored precipitate.

---

## CHAPTER VIII.

### Of Bismuth.

THIS metal is found native, and also in combination with sulphur, oxygen, and other metals. It is found plentifully in Saxony, and is met with in other parts of the continent and in Cornwall. It is brittle, shining, of a reddish-white color, and lamellated. Its specific gravity is a little below 10°. It melts at 480°, and its equivalent is 72°.

Bismuth may be obtained, for chemical purposes, by heating the oxyde to redness, along with charcoal.

If some bismuth be exposed, in a crucible, to a heat sufficient to melt it, in the open air, a crust soon collects on its surface, composed, chiefly, of the oxyde of bismuth. If this be exposed to a high temperature, in a furnace, the bismuth is volatilised and burns with a bluish-white flame. Exp.

Let a portion of the metal be melted in a crucible, and then set it aside to cool. When the surface has become solid, the greater part of the metal within being fluid, make a small hole in the middle of the surface, and invert the crucible, that all the liquid metal may flow out. On breaking into the hollow mass that is left, its internal surface will be found studded with crystals of metallic bismuth. The same thing may be accomplished, by melting the metal in an iron ladle; cool the bottom by dipping in cold water, before the crust is pierced, taking care that none of the water falls on the melted metal, an event that might be productive of injury to the operator, by causing a sudden and violent explosion. Exp.

*Oxyde of Bismuth.* There is but one oxyde of this metal, and that is composed of bismuth 72, one equivalent, and oxygen 8, one equivalent. It is of a yellow color, and can be procured by heating the subnitrate of bismuth to redness, so as to drive off all the nitric acid. At a full red heat, the oxyde is fused, and yields a transparent, yellow glass. At a higher temperature, it is sublimed. It unites with acids, and its salts, for the most part, are white. What is commonly called the white oxyde of bismuth, is not an oxyde, as we shall see in the sequel.

*Chloride of Bismuth.* If bismuth, in very fine powder, is dropped into chlorine gas, it takes fire, burns with a pale blue light, and is changed into a chloride, which was formerly called the *butter* of bismuth. It is composed of chlorine 36, and bismuth 72, or one equivalent of each.

This compound may be prepared, also, by heating two parts of corrosive sublimate with one of bismuth, and then driving over the excess of chlorine, and also the mercury, by heat. It is of a grayish-white color, opaque, and of a granular texture. It fuses at  $500^{\circ}$ , and bears a red heat, in close vessels, without being sublimed.

*Sulphuret of Bismuth.* This is a native production, although a compound, exactly similar, may be formed by fusing bismuth with sulphur. The component parts are, sulphur 16, and bismuth 72, or one equivalent of each. The dark colored precipitate, formed by passing a stream of sulphuretted hydrogen gas through a solution of a salt of bismuth, is a sulphuret of bismuth, of similar constitution with the native article.

*Nitrate of Bismuth.* If we add half an ounce (by weight) of powdered bismuth, to six drachms of nitric acid, (by measure,) diluted with half its bulk of water, and expose the mixture to a gentle heat, a smart action will ensue, in which part of the acid is decomposed, furnishing oxygen to the metal and changing it into an oxyde. The undecomposed acid joins the oxyde thus formed, deutoxyde of nitrogen being evolved during the process. On evaporating the solution by a gentle heat, and setting it aside to cool, crystals of nitrate of bismuth are deposited.

If a portion of this nitrate be thrown into pure water, there will be formed, instantly, a copious precipitate. This has been, generally, called the white oxyde of bismuth, and most of our Dispensatories and Pharmacopœias treat of the article under that title. But it will presently be shown, that it is not an oxyde, but a subnitrate. Formerly it was known as the *magistery of bismuth*, and on account of its whiteness, was employed as a cosmetic. It is, however, an inconvenient pigment, because it is instantly blackened by sulphuretted hydrogen. The reaction that ensues, on adding water to the nitrate, is made sufficiently plain, by the following diagram.

Before decomposition.		After decomposition.
134 or 1 equiv. of nit. of bismuth	{ nitric acid 54	188 bi-nit. bism.
	{ ox. bismuth 80	
134 or 1 equiv. of nit. of bismuth	{ nitric acid 54	
	{ ox. bismuth 80	
134 or 1 equiv. of nit. of bismuth	{ nitric acid 54	214 subnit. bism.
	{ ox. bismuth 80	
402	402	402

The resulting compounds are, an insoluble *subnitrate*, and a soluble *bi-nitrate*. The first is called a *subsalt*, because the metallic base is present, in the proportion of two equivalents to one of acid; it is such a compound as Berzelius calls a *di-nitrate*.

To show that the metal is actually contained in the supernatant liquid, as well as in the precipitate, decant it off and pass a stream of sulphuretted hydrogen gas through it. The appearance of a dark colored sulphuret of bismuth, thrown down in the form of precipitate, will be proof of the point in question.

Owing to the ready manner in which the nitrate is decomposed, and a white subnitrate thrown down, the solution of the nitrate, as it is first formed, is employed as a sympathetic ink.

Dip a clean pen into some of the newly made nitrate of bismuth, (not evaporated,) and write on a clean sheet of paper. When the writing is dry, no traces can be seen. If the paper be now dipt into water and immediately drawn out, the writing will be quite legible; and it will be seen, that every letter is marked with a white precipitate, detached from the nitrate by the decomposing power of water. Exp.

Take a quantity of the solution of nitrate of bismuth, and mix it with a solution of tartrate of potash, adding the latter as long as any precipitation takes place. By a double decomposition, two new salts will be produced, viz. the *nitrate of potash* and the *tartrate of bismuth*. This latter article was also used as a cosmetic, but is liable to the same inconvenience with the subnitrate. To illustrate this position, draw a face on paper, and paint it well with the tartrate of bismuth; then direct upon it a stream of sulphuretted hydrogen gas, from a bladder, furnished with a tube and stop-cock. In an instant, the pearly white will be converted into a dingy black. It is related, that tricks, on this principle, have been occasionally played with fashionable ladies, to their no small mortification. Exp.

The name *pearl-white*\* has been applied to the white oxyde or subnitrate, as well as to the tartrate.

The subnitrate was once a very popular medicine in the treatment of dyspeptic affections, but its fame was merely the creature of fashion, for with the changes of the latter, the celebrity of the medicine almost entirely vanished.

Orfila affirms, that the subnitrate has been employed by the English bakers, to whiten their bread; but Christison does not credit the statement. Whenever it is supposed to be present in bread, or other articles, the substance should be calcined in a crucible, and the residuum acted on by nitric acid. The solution thus formed, ought next to be treated with sulphuretted hydrogen gas, and the decomposition of the nitrate should be attempted, by means of water, as already illustrated.

## CHAPTER IX.

Of Manganese.

METALLIC manganese is not easily obtained, as it has a very great affinity for oxygen; it has been procured in small grains, by exposing the peroxyde mixed with charcoal, to the greatest

\* Called also *flake-white* and *Spanish-white*.

heat of a smith's forge. The sides of the crucible must be lined, previously, with a stiff paste made of charcoal, otherwise part of the oxyde will combine with the earthy matter, and fuse it into a glass, after part of its oxygen has been expelled.

It was by such a process, that Gahn and Scheele discovered this metal, in the year 1774. It is hard, brittle, of a grayish-white color, and granular texture. Its specific gravity is 6.8, and its equivalent 28. It is one of the most infusible metals, and requires a heat of  $160^{\circ}$ , Wedgewood, to melt it. Exposure to the air causes it to be tarnished, and when acted on by a red heat in that exposure, it absorbs oxygen rapidly. It is one of the metals that decompose water, with evolution of hydrogen gas, at ordinary temperatures; but the process goes on very tardily. The decomposition of that fluid is effected with more rapidity, at a red heat, and the metal is changed to the protoxyde.

It is on account of the very powerful attraction between manganese and oxygen, that this metal is never found in the earth, in a pure, metallic state, but always as an oxyde.

*Oxydes of Manganese.* Many writers name four oxydes of manganese; as, for instance, the *protoxyde*, *deutoxyde*, *red oxyde*, and *peroxyde*. But, as the *red oxyde* is, most probably, a compound of two others, it seems to be proper to limit the number of oxydes to three, the composition of which may be stated thus.

	Manganese.	Oxygen.	
Protoxyde	28 or 1 equiv.	+ 8 or 1 equiv.	= 36
Deutoxyde	28 or 1 equiv.	+ 12 or $1\frac{1}{2}$ equiv.	= 36
Peroxyde	28 or 1 equiv.	+ 16 or 2 equiv.	= 44

In addition to the above oxydes, there are two compounds of oxygen and manganese, which have been called *manganetic* and *manganous acids*.

The *protoxyde* may be obtained, by mixing any of the other oxydes of manganese, intimately, with charcoal, and exposing the mixture to a white heat, in a furnace. The carbon combines with part of the oxygen, and protoxyde of manganese is left. It may also be formed, by passing a stream of hydrogen gas over any of the oxydes of manganese at a bright red heat, having first placed them in a gun-barrel or porcelain tube, made to traverse the furnace, as in the process for making hydrogen by the decomposition of water. One end of the tube must be connected with an apparatus which can supply a constant stream of hydrogen gas, and to the other end another tube must be attached, that dips under water, so as to carry off the superabundant hydrogen gas. The excess of oxygen in the oxyde, unites with part of the hydrogen, forming water, while the protoxyde remains in the tube, in the form of green powder.



The protoxyde unites readily with acids, without any effervescence, producing precisely such salts as result from the action of the same acids on the carbonate of manganese. When quite pure, the protoxyde is entirely dissolved in cold, diluted sulphuric acid, and a colorless solution is formed. So entirely opposite are the accounts of eminent chemists, respecting the action of air and heat on the protoxyde, that one might suppose they wrote before, rather than after their experiments were made. Thus, one affirms that it attracts oxygen from the air, *rapidly*; another, that after an exposure during nineteen days, there was not the slightest change, either in weight or appearance. When heated, it takes fire, according to the testimony of some, while others were unable to perceive any such effect.\*

The *deutoxyde of manganese* occurs nearly pure in nature, and has been found as a hydrate, in fine, large crystals. We may obtain it artificially, by exposing the peroxyde to heat in an iron bottle; oxygen gas is evolved, and the deutoxyde remains of a dark brown color. As its proportion of oxygen is one and a half, some chemists have supposed it to be a compound of one equivalent of the protoxyde and one of the peroxyde, in order to get rid of the fraction.

The *peroxyde of manganese* occurs largely in the mineral kingdom, and from it the other preparations of manganese are usually obtained. Its uses in the formation of oxygen, chlorine, and in several other processes, have been already described.

This substance is the *black oxyde of manganese*, which was so useful to Scheele in some of his operations, and the nature of which he developed in 1774. It generally occurs in the massive form, has an earthy appearance, and is mixed with other substances, as silicious and aluminous earths, oxyde of iron, and carbonate of lime. It is found in very many locations; recently a vein of considerable size and tolerably good quality, was found within five miles of Philadelphia.

This oxyde may be prepared, by exposing the nitrate of manganese to a moderate red heat, until all the nitric acid is driven off. If the heat be too great, the oxygen of the peroxyde will also be expelled, in part.

Peroxyde of manganese is unaffected by exposure to the air. It is insoluble in water, and will not combine with acids nor alkalis. When boiled with sulphuric acid, it yields oxygen gas, and is thus changed to the protoxyde, which unites with the sulphuric acid, giving rise to the sulphate of manganese.

If we use muriatic acid, instead of the sulphuric, a muriate

\* See Turner's Chemistry.

of the protoxyde is formed, and chlorine is evolved. This will be seen by the following diagram.

Before decomposition.		After decomposition.	
37 or 1 equiv.	{ chlorine	36	36 chlorine.
mur. acid	{ hydrogen	1	9 water.
37 or 1 equiv.	mur. acid	37	
44 or 1 equiv.	{ oxygen	8	
peroxyde	{ oxygen	8	
of mang.	{ manganese	28	73 muriate of mang.
<hr/> 118		<hr/> 118	
<hr/>		<hr/>	

In this representation, two equivalents of muriatic acid act on one of the peroxyde. One equivalent of the acid is decomposed, as is also the peroxyde of manganese; the hydrogen of the decomposed acid joins one equivalent of oxygen, furnished by the peroxyde, and water is the product, the chlorine of the acid escaping at the same time in form of gas. The peroxyde being thus reduced to the protoxyde, unites with the equivalent of undecomposed acid, and the *muriate of the protoxyde of manganese* is obtained.

The following is the composition of *manganetic* and *manganeseous acids*, as stated by Reid.

	Oxygen.	Manganese.
Manganeseous acid	24 or 3 equiv.	+ 28 or 1 equiv. = 52
Manganetic acid	32 or 4 equiv.	+ 28 or 1 equiv. = 60

These acids are, comparatively, of recent discovery; and, even at the present time, there is not a little obscurity connected with their history.

If we take the black oxyde (peroxyde) of manganese, and mix it with an equal weight of nitrate or carbonate of potash, and expose the mixture to a red heat, a green colored, fused mass is obtained, which has long been known as the *mineral chameleon*, from the fact of its solutions changing color several times in a short period. This mineral chameleon, while it excited curiosity and speculation, remained unexplained, until the fact was announced, that manganese was susceptible of acidification, and that the acid, thus formed, united with the alkaline base, giving rise to a true chemical compound.

During the winter of 1831-2, I made a large quantity of this substance, without intending or expecting to do so; and as the same contingencies may fall to the lot of others, the incident is worth relating. I happened to be in the act of preparing oxygen gas from nitre, for my lectures on that subject, and was greatly surprised to find that the gas ceased to come

over much earlier than on former occasions. I was using the ordinary iron bottle and a bent gun-barrel, to which a leaden tube, twenty feet long, was attached. About a pound and a half of powdered nitre had been placed in the bottle, (capable of holding a half gallon by measure,) and I had not collected one hundred cubic inches of gas. I directed my assistant to remove the luting, and detach the leaden pipe from the gun-barrel, in order to ascertain, if possible, the source of disappointment. He informed me, at once, that the barrel and pipe were quite choked, with something like verdigris; and, at first sight, his impression seemed to be correct. The passage was almost entirely obstructed, and it is really wonderful, that an explosion had not taken place, for the bottle had been, for more than a half hour, at a full red heat. All at once, however, it occurred, that I had attempted to make oxygen gas in the previous winter, in the same bottle, and that I had then used the black oxyde of manganese, but had not taken the trouble to clean it after the operation was over.\* Of course, the bottle, when about to be placed in the furnace in the expectation of getting the pure gas from nitre, and in large quantities, contained, not simply the nitrate of potash, but the black oxyde of manganese, also. Here were the very articles requisite to the formation of the chameleon mineral, and my furnace was precisely the place for effecting the requisite combination.

It will be remembered, that the attempt, one year previous, to get oxygen from the black oxyde, failed. Possibly, the heat was not strong enough to decompose that oxyde. But we know, that an oxyde under these very circumstances, may absorb a larger portion of oxygen, so as to be oxydated to the maximum, or to be acidified; and this is precisely what happened in the case before us.

The decomposition of nitre always furnishes more or less oxygen, and consequently, in the early stage of the process, some portions of that gas were collected. But the manganese was brought into a condition, at the same time, which enabled it to take up the greater part of the oxygen, furnished by the decomposition of the nitre. It was changed into the *manganeseous acid*, and this, uniting with the potash of the nitre, formed the *manganesite of potash*, or the chameleon mineral;† which being for the most part condensed into a solid mass, filled the gun-barrel and tube, to the length of several feet. In ad-

\* This took place in the vicinity of Philadelphia, and from some cause or other, not known, the black oxyde did not afford oxygen in sufficient quantity for the purpose of experiment.

† This compound is deliquescent, and should, therefore, be kept in close vessels.

dition to the green, solid mass taken from these cavities, which was exceedingly hard, I collected several ounces of a most beautiful deep-green fluid; and in cleaning out the bottle, a quart of green, pulverulent matter was obtained.\*

The *manganetic acid* is said, by Dr Forchhammer,† to be produced in the solutions of the chameleon mineral. As these are passing from one hue to another, until the fluid is changed from green to red, the manganeseous acid of the compound, absorbs oxygen, and is converted into manganetic acid. This statement accords with the generally received explanation, and I do not see wherein it is defective.

Let ten or twelve grains of the powdered mineral chameleon be put into a deep ale-glass, or other vessel, and pour on it a little water. We shall very soon have a deep-green  
 Exp. solution, which, on adding more water, will pass through various shades of green, blue, and purple, and ultimately change to red. If the solution be now set aside for a few hours, it will be found that a reddish-brown flocculent matter, called the red oxyde of manganese, has subsided, and that the solution has become colorless.‡

As a proof that manganetic acid is thus formed, by the further acidification of a part of the manganeseous acid, the following experiment may be made.

To a portion of the solution of the chameleon mineral that has changed from green to red, add small quantities of  
 Exp. potash, and the color will retrograde, slowly, until the original green is perceptible. It would seem, that the red color depended on the formation of a small quantity of manganetic acid, and that this being neutralised by the addition of more potash, the original green of the unchanged manganite of potash was restored.

If the red solution be quickly evaporated, small prismatic crystals of a purple color are obtained, which prove to be the neutral manganate of potash. This salt yields oxygen to combustible bodies very readily, and with phosphorus, it detonates powerfully. In solution, it experiences decomposition from slight causes, and is converted into the red or brownish-red oxyde of manganese.

We have already made mention of the muriate of manganese, and now recur to it to state the process directed by Mr Faraday, to obtain it free of foreign matters.

\* Since I wrote the above, I have met with a very interesting paper on the manganite of potash, by Mitscherlich, in the *Annales de Chimie et de Physique*, for February, 1832.

† Annals of Philosophy, vol. xvi.

‡ These changes occur more rapidly if *hot water* be employed.

The peroxyde of manganese, being usually mixed with oxyde of iron, which is dissolved along with the oxyde of manganese when digested in sulphuric or muriatic acid, the process of Mr Faraday, will be found convenient when a pure solution of manganese is required. Mix one part of muriate of ammonia, in fine powder, with twice its weight of the peroxyde of manganese, and expose the mixture to a dull, red heat, in a crucible, for a quarter of an hour; the ammonia is disengaged, the oxyde loses oxygen, and the chlorine or the muriatic acid combines with metallic manganese alone, having a much greater affinity for this metal, at a red heat, than for iron. As great excess of peroxyde is used, no chloride of iron is formed; and a solution of the muriate of manganese is obtained, by digesting the remaining mass in water, and passing the solution through a filter.

To show the actual difference between the muriatic solution, prepared according to Mr Faraday's direction, and that made in the usual way, and which always contains more or less of iron, the following experiments may be resorted to. Add a few drops of the ferrocyanate of potash to the first solution, and a white precipitate of ferrocyanate of manganese will be thrown down, immediately. But if the same test be Exp. dropped into the ordinary solution, a rich, blue precipitate will appear, which indicates the ferrocyanate of iron.

A solution of the muriate of manganese has been used, for the purpose of estimating the strength of a solution of the chloride of lime. It may be easily obtained, by evaporating to dryness, the solution that remains after the preparation of chlorine, from common muriatic acid and peroxyde of manganese, then digesting the residuum in water, and filtering.

*Chloride of manganese* may be prepared, by evaporating a solution of the muriate to dryness by a gentle heat, and then exposing the residue to a red heat in a glass tube, at the same time passing a stream of muriatic acid gas through the tube. When the process is over and the apparatus has become cold, a pink-colored mass is discovered, which is deliquescent and of course soluble in water, being changed by that fluid into muriate of manganese. It is composed of 36 chlorine, and 28 manganese, or one equivalent of each component part.

*Fluoride of Manganese.* This compound has been recently noticed by Dumas and Wohler,\* who have the credit of its discovery. It is a greenish-yellow gas or vapor, somewhat of the hue of chlorine. It is formed, by mixing chameleon mineral with half its weight of fluor spar, and decomposing the mixture in a platina vessel, by adding strong, fuming sulphuric acid.

\* Edinburgh Journal of Science.



The calcium of the fluor spar (fluoride of calcium) is oxydated by the oxygen of a part of the manganeseous acid, and the manganese, thus detached, unites with the fluorine of the fluor spar, and so gives rise to the fluoride of manganese.

When this gaseous compound is mixed with atmospheric air, it instantly acquires a beautiful, purple-red color; and it is freely absorbed by water, making a solution which has the same red tint. It acts, instantly, on glass, forming fluosilic acid gas, and depositing a brown matter, that gives a deep purple hue to water.

The action of water on this gaseous matter, gives rise to the hydrofluoric and manganetic acids; and the same effect is produced by atmospheric air, on account of the moisture it contains. The water is decomposed, its hydrogen joining the fluorine to form hydrofluoric acid; while its oxygen unites with the manganese to form manganetic acid. The formation of the latter acid explains the red or purple color, above mentioned, and the action of the compound on glass, proves the presence of fluoric or hydrofluoric acid. The composition of this gas is, one equivalent of manganese, or 28, and four equivalents of fluorine, or 64.

The books notice but one compound of sulphur and manganese, which is denoted with sufficient plainness, by the term *sulphuret of manganese*.

This compound consists of one equivalent of each of its component parts, and its combining number is, consequently, 44. It may be made, by decomposing the sulphate of manganese, by means of charcoal, and also by passing a stream of sulphuretted hydrogen over the protosulphate, at a red heat. It yields very pure sulphuretted hydrogen gas, when acted on by diluted sulphuric or muriatic acid. The rationale corresponds with that given, for the formation of this gas from the sulphuret of iron or antimony.

---

## CHAPTER X.

### Of Chrome.

THIS metal has been so called, from a Greek word, signifying *color*, as indicative of its remarkable tendency to produce colored compounds. *Chromium* or *chrome* may be used, as equally significant of the object.

Vauquelin is the reputed discoverer of this metallic body,

and he found it in the year 1797, in a red mineral of great beauty, which proved to be chromate of lead. Owing to the intense affinity between this metal and oxygen, it is very difficult to exhibit it in the detached form, as a metal. It has been obtained, however, in minute portions, by heating, in a smith's forge, a mixture of charcoal and oxyde of chrome.

The color of this metal is white, or rather a yellowish-white, and it has a distinct metallic lustre. Its specific gravity is 6, and its equivalent 28. It is brittle, infusible, and resists the action of many acids, being slightly affected, even by the nitromuriatic acid. When it is exposed to heat, in connexion with the nitrate of potash, it is oxydised at the expense of that salt, and chromic acid is the result. It is therefore one of the metals, which admit, not only of oxydation, but of actual acidification.

Different views have been given of the compounds formed by chrome and oxygen, more especially as respects their number; some writers noticing but two, while others describe three. Reid adopts the larger number, as do also Berzelius, Ure, and others. Their composition appears to be as follows.

	Oxygen.		Chrome.	
Protoxyde	8	+	28	= 36
Deutoxyde	16	+	28	= 44
Peroxyde (chromic acid)	24	+	28	= 52

The *protoxyde* is easily formed, by exposing the salt, called chromate of mercury, to a strong red heat for half an hour. The oxyde of mercury in the salt, is decomposed and volatilised, and the chromic acid, parting with a portion of its oxygen, is reduced to the state of an oxyde, viz. the *protoxyde*. This has a fine green color, and is generally called the *green oxyde*; the solutions of such of its salts as are soluble, have the same color. We have, in this instance, a striking illustration of a law of chemical action, formerly mentioned, viz. that bodies in entering into combination, experience an entire change of color, as well as of other properties. In the present case, a metal, almost white, is changed into a green compound, by simple union with a colorless gas.

The *protoxyde* of chrome is infusible, indecomposable by heat alone, and is not acted on by oxygen or common air. Heated to dull redness with half its weight of potassium or sodium, a brown matter is produced, which, when cooled and exposed to the free air, burns with flame, and is changed into chromate of potash or soda, having a canary-yellow color. When deflagrated with nitre, it is oxydised to the maximum, or in other words, transformed into chromic acid. Fused with

borax, it imparts a beautiful green color, and this property not only affords an excellent test of its presence, but makes it exceedingly useful in the arts. The emerald owes its fine color to this oxyde.

The protoxyde is a salifiable base, and all its salts have a green color. They may be procured, by adding to a boiling solution of chromate of potash in water, an equal quantity of alcohol and muriatic acid, in small portions at a time, until the red tint of chromic acid entirely disappears, and the liquid acquires a green color. On pouring an excess of pure ammonia into the solution thus formed, a pale green, bulky precipitate is thrown down, which is composed of one equivalent of protoxyde of chrome and twenty-six equivalents of water. This precipitate is, of course, a hydrate, and it is readily soluble in acids.

The *deutoxyde* is made, by exposing the protonitrate of chrome to heat, till the fumes of nitrous gas are all expelled. A brilliant, brown powder, insoluble in acids, and scarcely soluble in alkalis, constitutes the residuum. It is not applied to any useful purpose.

*Chromic acid*, may be obtained in solution, by digesting 126 parts of dry chromate of barytes with 49 parts of sulphuric acid, diluted with seven or eight times its bulk of water. The sulphuric acid combines with the barytes, forming the insoluble sulphate, which remains mixed with the liquid, and the chromic acid is dissolved, imparting to it a very deep, ruby color. By concentrating the solution, it may be obtained in small crystals. It parts with a portion of oxygen when exposed to a red heat, and even when its solution in water is boiled or left in contact, for some time, with acids and many substances that have an affinity for oxygen, (as alcohol and other liquids containing a considerable quantity of inflammable matter,) the carbonaceous matter or hydrogen, uniting with the excess of oxygen in the chromic acid, and the green oxyde being retained in solution by the other acid that is mixed with it.

This acid was first obtained from the red lead ore of Siberia, which contains chromate of lead. On mixing this with carbonate of potash, a double decomposition ensued, and chromate of potash was formed. On adding nitric acid to this chromate, nitrate of potash was obtained, and the chromic acid separated. It has also been procured from the chromate of iron, which is found in large quantities near Baltimore, (Maryland,) and also in small portions, in Chester county, (Pennsylvania).

To obtain pure chromic acid, it must be distilled from fluor spar, (fluat of lime,) chromate of lead, and strong sulphuric

acid, in a leaden retort; a gaseous mixture of chromic and fluoric acids is evolved, that is readily absorbed by water. This mixed gas yields a thick orange-colored smoke, and on coming in contact with common air, it deposits small, red crystals of chromic acid. Ammoniacal gas, introduced into this gas, contained in glass jars lined with resin, burns with explosion. Crystals of chromic acid are also decomposed in ammoniacal gas, with a flash of light, and are changed into protoxyde of chrome. This acid is important, only on account of its highly useful compounds, which will be noticed in due order.

*Fluochromic Acid Gas.* This is the compound gas, just spoken of, and which deposits chromic acid, when brought in contact with the air. It acts readily on glass, depositing chromic acid, and forming fluosilicic acid gas. It is absorbed by water, and the solution contains hydrofluoric and chromic acids. It may be regarded, either as a compound of the elements, fluorine and chrome, or of hydrofluoric and chromic acids.

*Chlorochromic acid* is easily prepared, by mixing intimately, one part of common salt with two and a half of the chromate of lead, and exposing the mixture to heat in a glass retort, containing about an equal quantity of sulphuric acid. Sulphates of soda and lead are left in the retort, and the mixed acids unite together and are disengaged in the gaseous form. It has a rich red color, and may be collected in glass jars, over the mercurial trough. The experiment may be made, easily, on the small scale, with a few grains of the mixture, in a test-tube. It is instantly decomposed by water, and yields a solution of muriatic and chromic acids. We may regard it, either as a compound of muriatic and chromic acids, or of chlorine and chrome.

*Salts of Chrome.* The salts of chrome are divided into two classes, viz. those in which the protoxyde acts the part of a salifiable base, and those in which some other salifiable base is combined with chromic acid; the latter are the most important.

*Chromate of potash* is prepared from a mixture of nitre and the common chrome iron ore, usually termed chromate of iron, but which is composed of the oxydes of chrome and iron, mixed with earthy matter. The ore must be reduced to a very fine powder, and mixed with an equal weight of nitre, exposing the mixture to a bright red heat in a crucible, for about half an hour, when a small quantity of materials, as one or two ounces, is used. The nitric acid is completely decomposed, converting the oxyde of chrome into chromic acid, and the protoxyde of iron into peroxyde. The chromic acid unites with the potash, forming chromate of potash, and on boiling what remains in the crucible in water and filtering through paper, a clear solu-

tion of the chromate of potash is procured, and the iron and earthy matter remain on the filter. The free potash which the solution always contains, must then be neutralised by nitric acid, concentrating it afterwards by evaporation, that the nitrate of potash may crystallise, after which, the remaining liquid will deposit crystals of the chromate of potash, when allowed to evaporate spontaneously. They have a lemon-yellow color, and are soluble in twice their weight of cold water and in a much smaller quantity of boiling water.

Chromate of potash has a cool, bitter, and disagreeable taste. It is insoluble in alcohol, has an alkaline reaction, (that is, its solution changes vegetable yellow colors to brown,) and hence has been regarded by some, as a *subsalt*. It is, however, a perfectly neutral salt, containing one equivalent of chromic acid, 52, and one of potash, 48. Its crystals are anhydrous.

The *bi-chromate of potash* is prepared, by adding sulphuric acid to a solution of the neutral chromate, sulphate of potash being formed, while the chromic acid, set at liberty, unites with another portion of the chromate of potash which is not decomposed. The solution, at the same time, assumes a deep red color, and red crystals of the bi-chromate are procured, by spontaneous evaporation. They are less soluble than the crystals of the neutral chromate of potash, requiring about ten parts of cold water for their solution.

Add a solution of carbonate of potash to a solution of the bi-chromate, till the excess of acid is neutralised; the liquid  
 Exp. will be changed to a yellow color. Then drop in sulphuric acid, to neutralise the potash just added, and the red bi-chromate will be restored.

Fill a small crucible nearly half full of the bi-chromate, and invert over it another crucible of larger size. Let the lower  
 Exp. crucible be exposed to a bright red heat, in a much larger crucible, placed in a furnace, for the space of half an hour; the excess of acid will be decomposed, giving rise to oxygen gas and protoxyde of chrome, which remains mixed with the neutral chromate, thus formed. Let the green mass in the crucible (which has a glittering appearance) be reduced to powder, and boiled for a few minutes with five or six times its weight of water, and then filter through paper. The green oxyde of chrome will remain on the filter, and the liquid that has passed through, will have a deep yellow color, owing to some chromate of potash dissolved in the water.

The bi-chromate of potash is manufactured largely in this country, and is extensively employed in the art of dyeing.

*Chromate of barytes* is precipitated, when a solution of chro-



mate of potash is added to a solution of muriate of barytes, and muriate of potash remains in solution. It consists of one equivalent of acid, 52, and one of barytes, 78.

*Chromate of lead* may be formed, by adding a solution of chromate or bi-chromate of potash, to a solution of acetate of lead, as long as any precipitation takes place. It has a fine yellow color, and is the pigment so largely employed under the name of *chrome yellow*. It is composed of one equivalent of acid, 52, and one of protoxyde of lead, 112.

If a solution of chromate of potash be boiled with carbonate of lead, first reduced to a very fine powder, the *subchromate of lead* is obtained, having a very rich red color. It is composed of two equivalents of protoxyde of lead, Exp. and one of chromic acid, and hence has been called *di-chromate of lead*, to intimate the presence of two equivalents of base, in connexion with one of acid.

*Chromate of mercury* is easily prepared, by adding a solution of the chromate of potash to a solution of the protonitrate of mercury, as long as any precipitation takes place; the nitrate of potash remains in solution.

## CHAPTER XI.

### Of Cobalt.

THIS metal is said to be a constant ingredient in meteoric iron, and is usually combined with arsenic. It is brittle, of a reddish-gray color, and has a weak metallic lustre. Its specific gravity is 8.5, and its equivalent 26. It fuses at a heat below the melting point of iron, or at about 130° of Wedgewood, and when slowly cooled, it crystallises. It is attracted by the magnet, and may be rendered permanently magnetic. It undergoes very little change in the air, but when heated in open vessels, it absorbs oxygen. It is acted upon with difficulty by sulphuric and muriatic acids, but is readily oxydised by nitric acid.

*Cobalt* may be prepared from *zaffre*, (which is an impure oxyde of this metal, and the residuum of one of the principal ores of cobalt,) after exposing it to heat, in order to drive off the sulphur and arsenic which is usually combined with it. The cobalt of commerce generally contains small portions of arsenic, iron, nickel, and copper.

Several processes have been proposed, for the preparation of metallic cobalt from zaffre. Any arsenious acid may be removed, by dissolving it in dilute nitric acid and passing a stream of sulphuretted hydrogen through the solution, which will precipitate sulphuret of arsenic. A plate of iron will remove any copper, and on filtering the liquid, evaporating it to dryness, and digesting the dry mass in water of ammonia, nothing is dissolved but the oxydes of cobalt and nickel. Expel any excess of ammonia from the clear liquid by heat, and add a solution of potash cautiously, which will precipitate the nickel; filter the liquid immediately and boil it, when the oxyde of cobalt is separated, from which metallic cobalt may be obtained, by exposing it to a strong heat in a crucible, previously mixing it with charcoal, to effect its deoxydation.

Metallic cobalt is seldom employed in the arts; the impure oxyde, fused with silica and potash, yields a fine colored glass, to which the name of *smalt* has been given. *Powder blue* is smalt ground fine, and is much employed by sign painters.

*Oxydes of Cobalt.* There are two oxydes of this metal, viz. the protoxyde and peroxyde.

The protoxyde may be prepared, by decomposing the carbonate of cobalt by heat, in a vessel from which the atmospheric air is excluded. The carbonic acid is driven off, and the protoxyde remains. It may be formed, also, by precipitation. If we add pure potash to a solution of a salt of cobalt, a blue hydrate is thrown down, which is the protoxyde. It gives a blue tint to borax, when fused with it, and is employed in the arts, to color porcelain, glass, and earthenware. It is composed of one equivalent of oxygen, 8, and one of cobalt, 26.

The peroxyde is of a black color, and is made, by exposing the blue protoxyde, formed by precipitation, to the action of the air, that it may absorb an additional quantity of oxygen. It does not unite with acids, as the protoxyde does. When digested in muriatic acid, a protomuriate is formed, and chlorine is evolved. If we place it in close vessels, and apply a strong heat, it parts with a portion of oxygen, and is changed to the protoxyde. It is said to consist of one equivalent and a half of oxygen, and one of cobalt.

*Muriate of cobalt* is the only salt of this metal, that appears to be of any importance. It is used as a *sympathetic ink*, and is among the best articles for that purpose. It may be prepared, very easily, by digesting cobalt in an equal weight of muriatic acid, mixed with half its bulk of nitric acid and two or three parts of water. Evaporate the liquid thus obtained, to dryness, and then dissolve the residuum in water.

This solution gives traces of blue, but if a quantity of common salt be added, equal in weight to the cobalt employed, the marks will have a very beautiful green color.

To use this sympathetic ink, write a letter or a line or two, on clean white paper, and let it dry perfectly. Nothing can be discovered on the paper, until the sheet is applied near Exp. to a gentle fire, when all the letters are seen of a blue or green color, according as the ink has been made; and the traces disappear, soon after the paper is laid in a cold situation.

The trunk and branches of a tree may be rudely sketched on paper with Indian ink, and the foliage added, by using Exp. the solution of cobalt. The leafless tree will present itself, while the paper is kept cool, but so soon as it is heated, there will be an agreeable appearance of verdure.

## CHAPTER XII.

### Of Tungsten.

THIS word is derived from the Swedish language and means a ponderous stone. The name *wolfram* is given to the native tungstate of iron and manganese.

Tungsten is a very hard, brittle metal, of a grayish-white color, brilliant and nearly infusible. Its specific gravity is 17.6, and its equivalent 96. At a red heat, it absorbs oxygen from the air and becomes brown, and is then known, as the *brown oxyde*, supposed to be the *deutoxyde*. This oxyde does not form salts, either with acids or alkalis; but when the metal is oxydated to its maximum, and changed into tungstic acid, we have an agent which can unite with alkaline bases, and produce salts.

For these reasons, some have doubted the propriety of classing tungsten with the metals, but it is, notwithstanding, usually placed among them, by common consent.

It has been found in nature, only in the state of an acid, combined with the oxydes of calcium, iron, manganese, and lead. It may be extracted from tungstic acid, by mixing this with charcoal and oil, so as to form a paste, and then exposing the mixture to a strong heat in a crucible.

*Tungstic acid* is of a yellow color, insoluble in water, and without action on litmus paper. It forms, with certain bases, the salts called *tungstates*, all of which are decomposable by the

stronger acids. When strongly heated in open vessels, it acquires a green color and changes to a blue when exposed to hydrogen gas, at a temperature of  $550^{\circ}$ .

All the tungstates are insoluble in water, excepting those of potash, soda, ammonia, and magnesia.

## CHAPTER XIII.

### Of Molybdenum.

THIS name is of Greek origin, the word whence it is derived meaning *lead*; and the metal was probably so called on account of its resemblance to lead, in point of weight. Cronstadt gave it this name, but Hielm was the first who produced the metal in a separate state, (in 1782). Having been procured in very small grains, its peculiar properties have not been well ascertained. It is said to be a brittle and infusible metal, of a white color, having a specific gravity of about 8. Its equivalent number is 48.

The grains of molybdenum are changed by heat into a white oxyde, which rises in brilliant needle-shaped flowers, like those of antimony.

There are three grades of oxydation mentioned by some of the writers, viz.

	Oxygen.	Molybd.
The protoxyde consisting of	8	$8 + 48 = 56$
Deutoxyde (or molybdous acid)	16	$16 + 48 = 64$
Peroxyde (or molybdic acid)	24	$24 + 48 = 72$

The peroxyde or *molybdic acid* is the only compound of this metal that has been much examined. It is a white powder, having a specific gravity of 3.4. It has a sharp, metallic taste, reddens litmus paper, and forms salts with the alkaline bases. It is very sparingly soluble in water, but the molybdates of potash, soda, and ammonia readily dissolve in that fluid; and the molybdic acid is precipitated from the solutions, by the stronger acids. When melted with borax, it communicates a bluish color; and paper dipped in its solution, assumes a beautiful blue, on exposure to the sun.

This acid forms a class of salts, called *molybdates*, which have not been applied to any very useful purpose.

If a dilute solution of recently prepared muriate of tin, be  
 Exp. precipitated by a dilute solution of molybdate of potash, a beautiful blue powder will be obtained, although the solutions employed are colorless.

## CHAPTER XIV.

## Of Uranium.

URANIUM of *uranite*, is a solid, brittle, and very brilliant metal, of a deep gray color. Its specific gravity is 9, and its combining or equivalent number is 208. It is scarcely softened by the most intense heat, but when thus exposed, it absorbs oxygen and is converted into a black oxyde. Klaproth discovered it in a mineral called *pechblende*, which is said, by Ure, to be a sulphuret. It exists as an oxyde in the *green mica*, and in *uranochre*.

It is very difficult of reduction, as the intense heat of  $170^{\circ}$ , Wedgewood, is required to enable charcoal to deoxydate the yellow oxyde, which is formed by adding potash to the nitrate of uranium.

There are two oxydes of this metal, viz. the protoxyde and peroxyde. The first contains 8 oxygen, and 208 metal; the last has 16 oxygen, united to 208 of uranium. This metal combines with vitrifiable substances, and gives them a brown or green color. On porcelain, aided by the usual flux, it produces an orange hue.

## CHAPTER XV.

## Of Columbium.

MR HATCHETT discovered this metal in 1801, and called it after *Christopher Columbus*, the discoverer of the new world. A black, mineral substance had been for a long while in the British museum, without exciting any particular notice, until Mr Hatchett detected in it the metal now under consideration; the mineral was originally brought from Massachusetts, in the United States of America. It was afterwards found by a Swedish chemist, in the *tantalite*, and on the supposition that a new metal was brought to light, he called it *tantalium*. Dr Wollaston examined these nominally different products, in 1809, and established the fact of their identity.

Columbium is a metal of a dark gray color, somewhat resembling iron, and when reduced to powder, it becomes almost



brown. It is so hard as to scratch glass, and it is not attacked by acids. Its specific gravity is nearly 6, and its equivalent 144. It combines with oxygen in but one proportion, and the compound is known by the names of the white oxyde of columbium, and *columbic acid*. It is composed of oxygen 8, and columbium 144.

Columbium exists in most of its ores, in the form of columbic acid, and if the ore be fused with four times its weight of carbonate of potash, a soluble columbate of potash is obtained, from which columbic acid may be precipitated, as a white hydrate, on adding some of the stronger acids.

From the columbic acid, the metal may be obtained by the deoxydating power of charcoal, at a very elevated temperature.

The hydrated columbic acid is tasteless and insoluble in water; but when placed on litmus paper, it communicates a red tinge. It combines readily with alkalis and forms salts, which are not precisely neutral.

*Chloride of columbium* may be formed, by heating the metal in chlorine gas. It takes fire and burns actively, yielding a yellow vapor, which condenses in the cold parts of the apparatus, in form of a yellowish-white powder. When water is added to it, there is a hissing noise and elevation of temperature, and columbic and muriatic acids are produced. The hydrogen of the water joins the chlorine, to form the latter acid, while its oxygen goes to the columbium, and forms the columbic acid.

*Sulphuret of columbium* is formed, by heating the metal almost to redness, in the vapor of sulphur, or by passing the fumes of sulphuret of carbon over columbic acid, in a porcelain tube, at a white heat. The sulphuret of carbon and the columbic acid, are both decomposed; the oxygen of the acid joins the carbon and forms carbonic oxyde, which is evolved, while the sulphur of the sulphuret unites with the columbium of the columbic acid, giving rise to sulphuret of columbium.

Some of the other simple, non-metallic bodies combine with this metal, and perhaps all of them; but they do not call for our notice at this time.

---

## CHAPTER XVI.

### Of Titanium.

THIS metal has a red color, resembling that of copper. It is infusible, very brittle, and so hard as to scratch a polished surface of rock crystal. As yet, it has been obtained, only in the

form of brilliant pellicles. Its specific gravity is about 5.3, and its equivalent 32.

This metal was first observed by Mr Gregor, in Cornwall, (England,) in a kind of ferruginous sand, in a valley in that vicinity. He supposed that he had discovered a new metal, but was not able to reduce it. It is found in nature, in the state of a gray oxyde, and often combined with silex and lime.

The pure metal has been extracted from the oxyde, by means of charcoal and a powerful heat. The protoxyde and peroxyde are named in the books; the former containing one equivalent, and the latter two equivalents of oxygen, to one of metal. The peroxyde is also called *titanic acid*, because it readily combines with alkalis.

Titanium is susceptible of combination with several of the simple, non-metallic bodies, as chlorine, sulphur, &c.

To the foregoing list of metals, we might add *cerium*, *cadmium*, *tellurium*, *pluranium*, *vanadium*, &c.; but our limits do not allow of the introduction of articles, that have so little importance attached to them.

## ORDER FIFTH.

OF METALS, WHOSE OXYDES CAN BE REDUCED BY EXPOSURE TO HEAT,  
WITHOUT THE AID OF INFLAMMABLE MATTER.

### CHAPTER I.

#### Of Mercury.

THIS metal has been called *argentum vivum* or *quicksilver*, on account of its mobility, which indicated to the ancients some inherent living principle. It was likewise denominated *hydrargyrum*, on account of its fluidity and silvery color. *Mercury*, according to mythologists, was the winged messenger of the gods, and the patron of thieves; and Paris has well observed, that its almost universal agency, has made it the 'chief resource of the worst of thieves, viz. quacks and nostrum-mongers.' It

is found native in several countries, but is more generally met with in combination with sulphur, in the state of cinnabar. It is fluid at the ordinary temperature, has a specific gravity of 13.5; and its equivalent is 200.\* It boils, and may be distilled at 650°, and at 40° below zero, it may be frozen; in which state, its specific gravity is 15.

Mercury, in its metallic state, exerts no action on the animal economy, other than that of a mechanical kind. The Germans administer water, in which mercury has been boiled, as a vermifuge; but, as not a particle of the metal is dissolved, the effect must be ideal.

Pure mercury is not changed by exposure to the cold air, even though moisture be present. If it contain other metals, the amalgam may be oxydated, and a film be formed on its surface. The *æthiops per se*, of Boerhaave, was a blackish powder, which was supposed to be an oxyde of mercury. It was formed by agitating the metal, for some hours, in a bottle half full of common air. But it is now generally believed, that whatever properties the *æthiops* may have possessed, resulted from the oxydation of foreign metallic particles, while the mercury was wholly unaffected.

The fact, that globules of mercury have been found in various parts of the body, especially in the bony tissues, has given rise to much speculation. This result has been witnessed, usually, after large quantities of the various preparations have been taken, and for a long while; and the inference is unavoidable, that the powers of the living system are able to decompose the various compounds of this metal.

To obtain mercury in a pure state, the following means are resorted to.

Mix four or five ounces of the native or prepared sulphuret of mercury, with an equal weight of lime or iron filings, throwing an additional quantity of lime or iron over the mixture, and expose it for an hour or two to a dull red heat in an iron retort, or in an iron bottle, with a bent gun-barrel or other iron tube, adapted to it. The sulphur combines with the iron or the calcium of the lime, and metallic mercury is disengaged, being slowly volatilised and condensing in drops in the iron tube, the open extremity of which should be put into water.

This process may be imitated on a still smaller scale, by mixing twenty or thirty grains of the sulphuret with as much iron

\* I am aware, that some persons think 100 is the equivalent of this metal.

filings, and exposing them to heat in a glass tube, holding the part containing the mixture in a horizontal position over a spirit lamp, or placing it on the top of a chauffer, and surrounding it with small pieces of charcoal. A chauffer, with a piece cut out at the top of one of the sides, is extremely convenient for performing a number of experiments like this, on the small scale; the tube resting on the edge and being supported within the chauffer by the fuel, and without, by a brick or piece of wood. The temperature can be easily regulated; if the mixture is to be exposed only to a moderate heat, it will be unnecessary to cover it with charcoal, but if it must be subjected to a high temperature, then the tube should be coated with a proper luting, completely covered with a mixture of small cinders and pieces of charcoal, and a chimney put over it, which may be prolonged, if necessary, by another tube of nearly the same diameter, fitted to the top. The size of the glass tube must correspond with the quantity of materials used; it need not be larger than the tube employed for the reduction of arsenic, or it may be an inch in diameter, and ten or twelve inches long. The mixture should not occupy more than one or two inches of the sealed end of the largest tubes; and should always be placed in an inclined position, so that any watery vapor which may be disengaged and condensed on the side, may not fall back upon the hot part of the tube, otherwise it will be broken.

When a very small quantity of materials is used, the metallic globules are not seen very distinctly at first, but if the matter that is sublimed be taken out, and rubbed gently on a dry plate with a piece of paper, the metallic globules will become apparent.

When the materials are not well mixed, or the heat incautiously applied, a good deal of the sulphuret is sublimed, without decomposition, and is very apt to obstruct the tube, condensing, principally, a little way beyond the part to which the heat is applied; and, as serious accidents might take place from the accumulation of vapor, while there is no opening by which it may escape, the student cannot be too cautious in attending to this circumstance, in making experiments of this kind, especially in narrow glass tubes; when the temperature, however, is sufficiently high to soften the glass, no danger need be apprehended, as the glass will then be slowly blown out by the vapor within, till some part of it gives way and allows it to escape.

An iron tube, closed at one end by welding, as a piece of a gun-barrel, or an earthen tube closed with some clay or plaster of Paris, may be used, instead of a tube made of glass.

In this process, supposing iron alone to be used, it combines with the sulphur of the sulphuret of mercury, forming sulphuret of iron, while metallic mercury is disengaged. The following diagram gives the most probable view of the atomic proportions, in which the materials act upon one another, though I am not aware that this has been minutely examined.

Before decomposition.		After decomposition.	
232 or 1 equiv. of	mercury 200	200	200 mercury.
bi-sulphuret	sulphur 16		
of mercury	sulphur 16		
28 or 1 equiv. of iron	28		44 sulphuret of iron.
28 or 1 equiv. of iron	28		44 sulphuret of iron.
288	288	288	

Metallic mercury being frequently adulterated with a considerable quantity of other metals, such as lead, tin, zinc, and bismuth, it will be necessary to state the characters by which it may be distinguished when pure, and the method of purifying it from any foreign matter with which it may be adulterated.

Pure mercury has a bright, white, metallic lustre, and appears extremely mobile when poured from one vessel to another, or thrown upon a level surface, the globules in the latter case being round, having a very high edge, and being easily divided into a number of smaller globules, all of which appear equally mobile. It does not tarnish on exposure to the air; no film collects on its surface when shaken in a bottle, and when exposed to heat, it is completely volatilised.

When mercury has acquired a crust of oxyde from the action of acid fumes, or a quantity of dust collected on its surface, it is purified easily, by folding a piece of writing paper into a cone, leaving a small aperture at the bottom, about the size of the point of a pin, or a little larger, and pouring the mercury into this cone, supported in a glass funnel. The pure mercury will pass through, in a very slender stream, and the greater portion of the dust and oxyde adhere to the sides of the cone. A small quantity of mercury always remains at the bottom of the cone; it should not be forced through and mixed with the rest, but set aside by itself, or with other portions of impure mercury. When a large quantity of mercury is to be filtered in this manner, fresh portions should be poured into the filter from time to time, before what has already been put in ceases to drop, otherwise part of the dust or oxyde will be forced through along with the mercury. The vessel, into which the mercury is received, should be perfectly dry, otherwise it may not appear so pure as



it really is, since moisture prevents the globules that first fall through from coalescing so easily together.

If the mercury be adulterated with any of the metals we have mentioned, it has not that bright, metallic appearance which pure mercury always presents; a film soon collects on its surface, and another reappears whenever it is removed. It is much less mobile than pure mercury; when a small quantity is thrown on a flat surface, it does not divide so readily into globules, and they are not so round, but have an irregular appearance, and their edges, instead of being high and prominent, are nearly on a level with the surface itself; and when it is very impure, it presents more the appearance of a soft solid.

To separate these, the usual process is to distil it in an iron bottle, with a bent iron tube adapted to it. The Edinburgh college directs the mercury to be mixed with a sixth part of its weight of iron filings, and the mercury is obtained in a purer state in this manner, than when it is distilled without any admixture. The effect of the iron in this process is not very well understood; it has no great affinity for any of the other metals with which mercury is usually adulterated. Perhaps it acts, principally, by allowing the mercury to be converted more easily into vapor, and in smaller quantities at a time, lessening the risk of any being carried over mechanically, during the ebullition, as we know that water and other liquids can be made to boil several degrees below their usual boiling point, by introducing some pieces of wire or other solid matter, and that thus a constant stream of vapor may be made to arise from them instead of the liquid entering, only occasionally, into a state of violent ebullition, and then ceasing, for a short time, to give any more vapor, till it is produced with the same violence as before. Advantage is taken of this fact in the distillation of sulphuric acid, and it is probable, that in the distillation of impure mercury, the other metals which are mixed with it may have a tendency to pass over, when the distillation is conducted without the assistance of the iron filings.

In all cases, the distillation should be conducted with a very gentle heat, and the materials should never fill more than a third of the retort or bottle in which it is carried on; the extremity of the tube should be made to dip under water, and should be taken out, when the last portions of mercury have passed over. For distilling large quantities of mercury, one of the iron bottles in which it is sold will do extremely well, fitting a bent gun-barrel accurately to it by grinding; it may be heated by an open fire, placing it on a piece of brick or on one or two bars of iron, laid across each other, to raise it an inch or two above the

grating, surrounding it about half way up with burning fuel, (a mixture of charcoal and cinders should be used,) and taking care to moderate the heat, whenever it begins to boil. Mr Faraday recommends copper filings to be mixed with the iron filings.

The iron filings are generally mixed with a little oil, which is decomposed during the distillation, and a small quantity of an empyreumatic oil passes over with the metallic mercury and condenses in the water, often preventing the globules from uniting together, so that they assume the appearance of a soft solid. When this takes place, the water should be poured off, and a small quantity of a solution of caustic potash poured over them; this removes the oil, and on washing them with water, they readily unite, after which they should be passed through a paper filter.

Though mercury is obtained sufficiently pure for ordinary experiments, by distilling it cautiously in the manner that has been described, it still frequently contains a small portion of zinc; this may be removed by shaking it with diluted nitric acid in a bottle, and then pouring both into a plate, where they may be left together for a few days, after which the mercury must be washed and filtered as before; the acid may be diluted with ten or twelve parts of water.

When only a small quantity of mercury is to be purified, as an ounce or a pound, Dr Priestley's method will be found most convenient. It consists merely, in shaking the mercury briskly in a bottle capable of containing four or five times as much, blowing into it occasionally with bellows, to renew the air, and continuing till a black matter gathers together, which may be easily separated from most of the metallic mercury by a paper funnel, after which it should be returned again, and the operation repeated till no more oxydation takes place, when the mercury will become extremely clean and mobile, the brightening taking place all at once, as the last portions of the other metals are oxydated.\*

Mercury is susceptible of two degrees of oxydation, and both of the oxydes are capable of forming salts, when combined with acids. They are constituted as follows.

Mercury.	Oxygen.
Protoxyde 200 or 1 equiv. +	8 = 208
Peroxyde 200 or 1 equiv. +	16 = 216

\* The refuse matter of this operation, is the *æthiops per se*, already noticed.

*Protoxyde of mercury*, called also *black* or *ash-colored oxyde of mercury*, is prepared, most easily, by mixing chloride of mercury with a solution of potash, in a mortar, rubbing them together for a quarter of an hour or twenty minutes, and taking care to have an excess of alkali, that the decomposition may be complete. For every 236 grains of the chloride employed, 100 grains of fused potash may be taken and dissolved in two or three ounces of water, allowing the solution to stand till it becomes clear, when it may be decanted for use. Muriate of potash remains in solution, and the protoxyde may be separated by filtration; washing it with cold water, and keeping it in a dark place, as it soon begins to be decomposed on exposure either to heat or light, one portion losing oxygen, which combines with another, so that small quantities of the metallic mercury and peroxyde of mercury are then found to be mixed with the protoxyde.

Instead of using a solution of potash, the Edinburgh and London colleges prepare their protoxyde from the chloride of mercury, (formerly called muriate of mercury,) by the agency of lime water. The same reaction takes place, as when potash is used; half an ounce of the chloride may be taken for every five pounds of lime water, boiling them for a quarter of an hour, after rubbing them together in a mortar, mixing a small quantity of the lime water at first with the dry powder, that it may be easily moistened. Muriate of lime remains in solution, and the oxyde must be washed on a filter with distilled water; the following diagram shows more precisely the nature of the reaction.

Before decomposition.			After decomposition.	
28 lime		28	65 muriate of lime.	
9 water	{	1 hydrogen		
		8 oxygen		
236 chloride		36 chlorine		
of merc.		200 mercury	208 protoxyde of mercury.	
<hr/> 273		<hr/> 273	<hr/> 273	

The protoxyde of mercury may be obtained, also, by adding a solution of potash or soda to a solution of the nitrate of mercury, the alkali uniting with the acid, while the oxyde is precipitated.

By triturating metallic mercury with manna, sugar, honey, and a number of other substances, the metallic globules disappear, and a mass is obtained, of a dark color; many consider the metallic mercury to be merely reduced to a very minute

state of division in this case, while others affirm that it is at the same time oxydated.\*

When the protoxyde of mercury is quite pure, it has a dark color, and is completely dissolved by acetic acid, but is quite insoluble in muriatic acid.

*Peroxyde of mercury* may be obtained, by dissolving three parts of metallic mercury in four of diluted nitrous acid, (made by mixing equal weights of the strong acid and water,) evaporating the solution to dryness, and then reducing it to powder and exposing it to a stronger heat in an evaporating bason over a good chauffer, till it assumes a deep red color. It should be covered with a flat glass plate, which allows the progress of the decomposition to be observed; a large quantity of ruddy fumes is disengaged, which disappear when all the powder has acquired a dark color, after which it must be removed from the fire, otherwise it will be resolved into metallic mercury and oxygen gas. As it cools, it assumes a bright red color, and the lowest portion is usually obtained in the form of brilliant scales; this appearance depends, probably, on the pressure of the superincumbent mass, as it is always seen more distinctly when a large quantity of peroxyde is prepared. It is usually called *red precipitate*.

In this process, the metallic mercury decomposes part of the nitric acid and is converted into peroxyde of mercury, which combines with the rest of the acid, so that the dry mass, which is obtained in the first stage of the process, is a nitrate of the peroxyde of mercury. The nitric acid is afterwards almost entirely expelled, being resolved by the heat into nitrous acid and oxygen gas. The quantity of nitric acid that remains in combination with the oxyde, is extremely small. Another method of preparing peroxyde of mercury, consists in exposing metallic mercury to the air, at a temperature between 500° and 600°, when it combines slowly with oxygen and is converted into red scales; these were formerly called *precipitate per se*, and are larger than those procured by the decomposition of the nitrate; they have not the same shining appearance, however, and not containing any nitric acid, they are not so acrid. As upwards of a fortnight is required to prepare a few grains of peroxyde

\* This compound is usually called the *blue pill mass*. Mr Carpenter, of Philadelphia, prepares the protoxyde in a pure form, and sells it as a substitute for the *pill mass*. One fourth of a grain, made into a pill with some conserve of roses, is equal to a grain of the *blue pill mass*, and the strength of the dose is more uniform. The *mercurial ointment* is supposed to contain mercury, in the form of protoxyde, but this, too, is a matter of controversy, and Carpenter's protoxyde, triturated with lard, is regarded, by many, as a better article than the ordinary ointment of the shops.

in this manner, it is a process that is seldom resorted to. In order to prevent the escape of the mercury, in form of vapor, a matrass, with a stem nearly three feet long, is selected. The aperture is sufficiently large, to admit the necessary supply of air.

Peroxyde of mercury is frequently adulterated with oxyde of lead, which may be easily detected, by exposing it to heat on charcoal before the blowpipe. If the peroxyde be pure, it will be completely dissipated, oxygen gas being disengaged and metallic mercury volatilised; if, however, any oxyde of lead should have been mixed with it, a globule of metallic lead will remain on the charcoal. If it be purchased in the form of scales, it is generally obtained perfectly pure, and any foreign admixture could easily be detected, by bare inspection; but when it has been reduced to powder, it may then be suspected to have been adulterated. The red oxyde, made with nitric acid, and usually called *red precipitate*, is employed in the practice of medicine as an escharotic, and generally in the form of an ointment. It is slightly soluble in water, but extensively so in nitric acid, without any effervescence.

The *precipitate per se* is acrid and caustic. It is soluble in several of the acids, without decomposition. It is also slightly soluble in water, and is precipitated from the solution, by liquid ammonia. It was a favorite remedy with the celebrated John Hunter; he gave it in quarter grain doses, combined with opium.

There are several tests by which mercury may be detected in solution. The most delicate, perhaps, is that proposed by Mr Sylvester. A drop of the liquid suspected to contain it, is to be placed on a piece of gold leaf, or any piece of solid gold, and the point of a nail or penknife or of any small piece of iron or zinc, placed in contact with the moistened surface; if any mercury be present, the gold will immediately become white where it is touched by the other metal, uniting with the mercury and forming a solid amalgam, which retains its white color after the fluid has been wiped off. Exp.

Put a piece of copper into a solution of any salt of mercury; part of the copper will be dissolved, combining with the oxygen of the oxyde of mercury and the acid with which it was previously united, while an equivalent quantity of metallic mercury will be precipitated. Exp.

Add a solution of potash to a solution of a salt of mercury; oxyde of mercury will be immediately precipitated. If the mercury in the liquid shall have been in the form of a protoxyde, the precipitate will be of a dark color, but when it contains



the peroxyde alone, the precipitate has, sometimes, a reddish color, but is generally yellow, the peroxyde disengaged combining with a portion of water and forming a hydrate. Collect the precipitate on a filter, dry it, and expose it to heat at the bottom of a small test-tube, over a spirit lamp, when globules of metallic mercury will be seen.

Put a small quantity of a solution of a salt of the protoxyde of mercury into a glass, and fill it up with lime water; Exp. protoxyde of mercury will be immediately thrown down and give a dark color to the liquid.

Into another glass, put a similar quantity of a solution of a salt of the peroxyde of mercury, and add lime water as Exp. before; the solution will become of a yellow color, from the separation of peroxyde of mercury in combination with water.

Add some hydrosulphuret of ammonia, or pass a stream of sulphuretted hydrogen gas through a diluted solution of Exp. a mercurial salt, and a copious black precipitate of sulphuret of mercury will be thrown down, probably in combination with water.

Add a solution of the muriate of tin (muriate of the protoxyde) to a solution of a salt of the protoxyde of mercury; a copious Exp. ash-gray colored precipitate will immediately appear, consisting principally of metallic mercury, the oxyde of tin combining with the oxygen with which the mercury was previously united, and passing to a higher state of oxydation.

A solution of the ferrocyanate of potash gives a white precipitate with salts of mercury; muriatic acid and solutions of the muriates give a white precipitate with salts Exp. of the protoxyde, which is composed of the chloride of mercury.

All the salts of mercury are completely decomposed or volatilised, by exposure to a dull red heat.

#### SALTS OF MERCURY, CHLORIDES, IODIDES, AND SULPHURETS OF MERCURY, ETC.

In considering the mercurial preparations, we shall depart from the order hitherto observed, when treating of other metallic compounds. Next to the oxydes, we have usually introduced the chlorides, but in the present case, we are obliged to defer the notice of the chlorides, until we shall have examined the nitrate and sulphate of mercury, as these are sometimes used in preparing the chlorides.

*Nitrate of mercury* (protonitrate) may be obtained, by digesting metallic mercury in one and a half times its weight of diluted

nitric acid, (prepared by mixing one part of acid with four of water,) allowing the mixture to evaporate spontaneously, and adding a small portion of metallic mercury, if the quantity employed should have been completely dissolved. The crystals that are deposited, are not completely dissolved by water, unless they contain a slight excess of acid.\*

*Pernitrate of mercury* is obtained, by heating metallic mercury in an excess of strong nitric acid, a large quantity of nitric oxyde gas being disengaged; it is deposited in transparent crystals as the liquid cools; if thrown into a large quantity of water, it is decomposed, a nitrate of the peroxyde with an excess of acid remaining in solution, while a subnitrate is precipitated, composed of two equivalents of the peroxyde and one of nitric acid.

In preparing nitrate of mercury for different purposes, great attention must be paid to the strength of the acid employed, the temperature to which the mixture is exposed, and the relative proportions in which the acid and the metal are mixed together, as all these circumstances have an important influence upon the oxydation of the mercury and the nature of the resulting salt. If the acid be strong, and a larger quantity employed than is necessary to dissolve the mercury, and if the solution be assisted by heat, the mercury always attracts two equivalents of oxygen, and a pernitrate of mercury is obtained; but when the acid is diluted with three or four parts of water, the solution allowed to go on at natural temperatures, or assisted only by a very gentle heat, and more mercury is used than the acid can dissolve, a protoxyde will be obtained, which unites with a smaller quantity of acid. Now this distinction is a matter of practical importance; and unless attended to, there can be no certainty, in attempting to form calomel from the nitrate, that the product will not be corrosive sublimate. This will be seen in the sequel.

*Sulphate of mercury* is prepared, by boiling two parts of metallic mercury to dryness, with two and a half of sulphuric acid, exposing the mixture to heat in a glass vessel, over a common fire. On the small scale, an ounce of mercury with the proper quantity of acid, may be boiled to dryness over a common fire or good chauffer, taking care to avoid the fumes that are evolved. The mixture should be boiled slowly, otherwise a quantity of the sulphate will be thrown out, and it should be removed from the fire whenever it is dry. It is obtained in the form of a white, crystalline powder, when well prepared, and does not

\* This salt constitutes the basis of the well known *citrin ointment*.

deliquesce on exposure to the air. The following diagram represents its composition and the theory of its formation; one portion of the sulphuric acid affords oxygen to the mercury, and sulphurous acid is disengaged, while another portion unites with the oxyde, and very pungent, suffocating fumes are evolved.

Before decomposition.		After decomposition.	
Sulphuric acid	{ sulphs. ac. 32	32	sulphurous acid.
	{ oxygen 8		
Sulphuric acid	{ sulphs. ac. 32	32	sulphurous acid.
	{ oxygen 8		
Sulphuric acid	40		
Sulphuric acid	40		
Mercury	200	296	persulph. mercury.
	<hr/> 360	<hr/> 360	

This diagram represents the action of four equivalents of acid on one of metal. Two equivalents of the acid are decomposed into sulphurous acid and oxygen, while two remain unaltered, for the purpose of combining with the peroxyde of mercury thus formed. Two equivalents of sulphuric acid, joined to one equivalent of peroxyde of mercury, furnish a compound, the correct name of which must be, *bi-sulphate of the peroxyde of mercury*. The principles of our nomenclature demand this appellation, but as the name is too long to be employed in the diagrams, we shall make use of the more common term, persulphate.

Mix 296 grains of the salt thus formed, with 200 of metallic mercury, and rub them intimately together in a mortar; the metallic mercury divides the oxygen and acid contained in that salt, with the 200 parts of mercury which it already contains, and 496 parts of *sulphate of the protoxyde of mercury* are obtained, the different materials arranging themselves in the manner represented in the annexed diagram.

Before decomposition.		After decomposition.	
200 mercury	200	248	sulph. merc.
296 persulph. of mercury	{ oxygen 8		
	{ sulph. acid 40		
	{ oxygen 8		
	{ sulph. acid 40		
	{ mercury 200	248	sulph. merc.
<hr/> 496	<hr/> 496	<hr/> 496	

Throw half an ounce or an ounce of the persulphate of mercury, heated to the temperature of  $400^{\circ}$  or  $500^{\circ}$ , into five or six pounds of boiling water, in a large glass flask or earthen bason. A yellow colored precipitate will be immediately thrown down, composed of one equivalent of sulphuric acid and one of the peroxyde of mercury, another portion of the peroxyde remaining in solution, with an excess of acid. I am not aware that the latter has been very accurately examined; the annexed diagram gives a precise view of the nature of the reaction, supposing the salt that remains in solution, to contain only one more equivalent of acid than the persulphate; the yellow colored precipitate is usually termed, *subsulphate of mercury*, or *turbith mineral*.

Before decomposition.			After decomposition.	
296 persulphate of mercury	{ sulph. acid	40	336 supersulph. merc.	
	{ sulph. acid	40		
	{ perox. merc.	216		
296 persulphate of mercury	{ sulph. acid	40	256 subsulph. merc.	
	{ sulph. acid	40		
	{ perox. merc.	216		
592		592		592

It should be washed repeatedly with water on a filter, and then set aside to dry.

I have here given the usual names for the products of this decomposition, but they are not strictly correct. The terms *sub* and *super* denote nothing, with sufficient accuracy, for the modern chemist, and ought to be laid aside. It is plain, that the *turbith mineral* is a compound of one equivalent of sulphuric acid, 40, and one equivalent of peroxyde of mercury, 216; it is, therefore, a *sulphate of peroxyde of mercury*. What is called *supersulphate* of mercury, is composed of three equivalents of sulphuric acid, 120, and one equivalent of peroxyde of mercury, 216; it is, therefore, really a *trisulphate of peroxyde of mercury*.

The turbith mineral was formerly a very popular emetic, especially with the French physicians, and is really a good article.

*Chlorides of Mercury.* These are two in number, usually called *calomel* and *corrosive sublimate*, but more correctly, *chloride* and *bi-chloride*.

The first is also sometimes styled *protochloride*, *submuriate*, and *mild muriate*. *Panacea of mercury*, *sweet sublimate*, *aquila alba*, *manna metallorum*, are among the old and obsolete names.

The bi-chloride has been called *oxymuriate of mercury*, *corrosive muriate*, *deutochloride*, &c.

Although not in strict accordance with our previous arrangements, we shall speak of the bi-chloride, before we examine the nature, properties, &c. of the chloride.

*Bi-chloride of mercury* (corrosive sublimate) is usually obtained, by exposing a mixture of one equivalent of bi-sulphate of the peroxyde of mercury to heat, along with two equivalents of chloride of sodium, and subliming. The fumes and watery vapor that are evolved, are loaded with bi-chloride of mercury, and must be carefully avoided, as they are highly deleterious. The sand bath, in which the sublimation is effected, should be directly under a chimney, or in the open air.

The following diagram is designed to convey a satisfactory view of the reaction that ensues.

Before decomposition.				After decomposition.	
296 or 1 equiv. persulph. of mercury	{	mercury	200	X	272 bi-chloride merc.
		oxygen	8		
		sulph. acid	40		
		oxygen	8		
120 or 2 equiv. chloride of sodium	{	sulph. acid	40		72 sulphate of soda.
		chlorine	36		
		chlorine	36		
		sodium	24		
		sodium	24		72 sulphate of soda.
416			416		416

The 272 parts of bi-chloride are sublimed and condensed in a crystalline cake, lined with a beautiful net-work of prismatic crystals, and the 144 parts (or two equivalents) of sulphate of soda remain at the bottom of the vessel. The heat should be moderate, so as to avoid loss and risk, from the evaporation of the bi-chloride.

For conducting this and similar sublimations, we may use a glass cucurbite, an earthen bottle, or a common glass bottle, coated with clay and sand; or a common oil flask may be taken for small quantities, and will answer very well. The vessel should be one-third full of the mixture to be sublimed, and the sand should reach no higher up than the surface of the mixture.

If the bi-chloride be pure, it will be entirely volatilised by heat, and may be easily distinguished from calomel, by its solubility in water, and the yellow precipitate which it gives with lime water.

The bi-chloride is of a white color, and but little changed by exposure to the air. Its taste is acrid, astringent, and metallic.



Its specific gravity is 5.2. It was formerly called *oxymuriate*, because it was supposed, that the muriate of soda (chloride of sodium) being decomposed, furnished muriatic acid, which combined with oxygen, furnished by the oxysulphate of mercury, and so gave rise to oxymuriatic acid. This latter joining the mercury, gave the product, which, for many years, was called oxymuriate of mercury.

It is evident, from the diagrams given, and from analysis, that the bi-chloride consists of two equivalents of chlorine, united to one equivalent of mercury. *Deutochloride* expresses the same fact, but we give a preference to the prefix *bi*.

This preparation is soluble, compared with the chloride. It requires eleven parts of cold water, three of boiling water, and four of alcohol, for solution. It is also very soluble in ether, and it is affirmed, that this liquid has the curious property of abstracting it from the watery solution, when the mixture is freely agitated. As light partially decomposes it, opaque bottles should be used for its preservation.

The incompatibles of corrosive sublimate are, the *carbonates of the fixed alkalis*, which precipitate it of a yellow color; *ammonia* or the volatile alkali, which throws down a white precipitate; *lime water*, which occasions a deep-yellow precipitate, constituting the *aqua phagedenica*;\* also *tartar emetic*, *nitrate of silver*, *acetate of lead*, *sulphur*, *sulphuret of potash*, and the *soaps*. *Iron*, *lead*, *copper*, *bismuth*, and *zinc*, in the pure metallic state, also decompose it, forming precipitates which consist of an amalgam, in most cases; hence, glass and earthen mortars should be used for preparing prescriptions of this article, and not such as are made of metal.

In addition to the above incompatibles, it is known, that the following vegetable infusions occasion precipitates, when corrosive sublimate is added to them, viz. the infusions and decoctions of *chamomile*, *horse-radish root*, *colombo*, *catechu*, *cinchona*, *rhubarb*, *senna*, *oak bark*, &c.

Corrosive sublimate is employed in the practice of medicine, internally and externally. It constitutes the basis of almost all the panaceas, and other nostrums, for the cure of syphilitic and scrophulous affections; indeed it is not unsafe to affirm, that every quack remedy for these diseases, that has obtained any thing like celebrity, has been shown to contain more or less of corrosive sublimate.†

\* To form this article, which was once a popular lotion for ulcers, two grains of corrosive sublimate are added to two ounces of lime water.

† Corrosive sublimate dissolved in alcohol, is a good article to destroy bed-bugs.

But this article is the occasion of much mischief, when taken in an overdose, either accidentally or by design. It is one of the most virulent poisons that the mineral kingdom furnishes, and on that account, merits special notice.

For an account of the symptoms induced by this poison, we refer to Christison, Orfila, and other writers. As respects the treatment, so far as it is chemical, we shall offer briefly, what appear to be the most important facts.

Orfila ascertained, by direct experiment, that albumen had the power of decomposing corrosive sublimate, and throwing down calomel; of course, it must detach one equivalent of chlorine from the bi-chloride, and form the chloride, which is destitute of poisonous properties. The best remedy, therefore, because usually at hand, or easily obtained, is the white of eggs. Let two or three dozen of eggs be had, and separate the whites; to these add cold water enough to suspend them, and then give the patient as much of the liquor as possible.

More recently, it has been discovered, that vegetable *gluten* has the power of decomposing corrosive sublimate; and to this property we referred in our remarks on gluten. To administer this remedy, without delay, we may give wheat flour, suspended in water, in large quantities, and often repeated. If neither albuminous nor glutinous matter can be procured, milk and water may be resorted to.

In addition to what has been said on detecting corrosive sublimate, we add the following expedient of Brugnatelli, to distinguish *corrosive sublimate* from *arsenic*. 'Take a quantity of fresh wheat starch, mix it with the water, and add a sufficient quantity of iodine, to give the liquid a blue color; if *corrosive sublimate* or *arsenic* be added to this liquor, the color will be alike destroyed, and a red color will appear. If this change has been effected by arsenic, a few drops of sulphuric acid will restore the blue color; but if it was produced by the corrosive sublimate, the addition of the acid will cause no alteration.'

It is proper for me to say a word, respecting the mode of writing this article in prescriptions. There can be no doubt as to its correct chemical name, agreeably to the principles of modern nomenclature; but as every one knows what is intended by corrosive sublimate, and as mistake might occur in the use of the other terms, it is best to employ the old name, for the present. Possibly, at some period, not very remote, the improved state of practical pharmacy may be such, as to justify the exclusive use of the terms that are in strict accordance with the principles of modern chemistry.

*Chloride of mercury, or calomel*, is one of the most important

compounds with which chemistry has any concern. It was called a muriate, because it was supposed, that when the mixture of sulphate of mercury and muriate of soda was sublimed, a double decomposition occurred, by which a real muriate was produced. The reason for calling it chloride or protochloride, will be obvious to the attentive reader of the sequel. As respects the origin of the word *calomel*, there is some diversity of opinion. The meaning of the word itself, is *fair black*. Some have supposed the name was given, on account of the change from black to white, experienced by the mixture during its sublimation. Others have sought a solution, in the common operation of calomel, whereby it has been affirmed to be a *good* or *fair* remedy for *black bile*.

We prefer the name *calomel* for common use, to that of chloride, notwithstanding the chemical accuracy of the latter. It is one that is familiar to almost all persons; and from the use of which, no evil can arise. The terms muriate and oxymuriate, were calculated to mislead the young apothecary, and his difficulty was increased, by the introduction of the chloride, bi-chloride, proto and deutochloride.

Calomel, or chloride of mercury, may be prepared from the bi-chloride, the sulphate or nitrate of mercury.

When the bi-chloride is employed for this purpose, every equivalent must be mixed with an equivalent of mercury, rubbing them together in a Wedgewood mortar, until the globules of the metal completely disappear, and the mixture is converted into an ash-gray colored powder. It should be turned over from time to time with a spatula, when it appears to cake together at the bottom of the mortar, and great care must be taken to avoid the small particles of the bi-chloride, that rise as it is reduced to powder, as they are extremely deleterious; a few drops of water added to the mixture, prevent these particles from being disengaged, and facilitate the action of the mercury on the bi-chloride.

The gray colored mass must then be sublimed in a cucurbite, a glass vessel with a capital that is accurately fitted to it by grinding; a small tube proceeding from it, carries off any watery vapor that may be condensed. It is heated most conveniently, by placing it in a sand bath, the lower part of the body of the cucurbite being covered by sand, and the calomel condensing in the upper part as it is sublimed; the mixture should never occupy more than a third of the cucurbite. On a smaller scale, the sublimation may be easily effected in a glass tube. The following diagram explains the theory of the action.

Before decomposition		After decomposition.	
272 bi-chloride of mercury	{ mercury 200	236 chloride of merc.	
	{ chlorine 36		
	{ chlorine 36	236 chloride of merc.	
200 mercury	200		
<hr/> 472	<hr/> 472	<hr/> 472	<hr/>

A small quantity of the bi-chloride of mercury is always formed during the decomposition, and a little metallic mercury also appears, generally, between the sides of the cucurbit and the crystalline cake of calomel. By repeating the sublimation, it is usually obtained in a pure state, or the metallic mercury may be separated mechanically, and the bi-chloride removed, by reducing the cake to a fine powder and digesting it with water, which dissolves the bi-chloride, but has no action on calomel. The London college directs a solution of the muriate of ammonia to be used for this purpose instead of water, the bi-chloride being much more soluble in it than in water.

If calomel be suspected to contain bi-chloride of mercury, a small quantity should be boiled in distilled water for a few minutes, or in a solution of the muriate of ammonia, and the liquid filtered. If any bi-chloride of mercury be mixed with the calomel, it will be dissolved, and on adding a solution of potash to the clear liquid, peroxyde of mercury will be precipitated of an orange color, if water alone has been employed; but the precipitate will be of a white color, if the calomel has been boiled in a solution of the muriate of ammonia.

The next method of preparing chloride of mercury, is from a mixture of the sulphate of mercury (sulphate of the protoxyde) and common salt, (chloride of sodium). The materials must be intimately mixed together, and exposed to heat in the manner we have already described. Every 248 parts of the sulphate (one equivalent) require 60 parts of chloride of sodium, (one equivalent,) sulphate of soda remaining in the bottom of the cucurbit, while chloride of mercury is sublimed as before. The following diagram gives a more precise view of the nature of the reaction.



Before decomposition.		After decomposition.	
248 sulph. of merc.	{ 208 protox. of mercury sulphuric acid chlorine sodium	{ merc. 200	{ 236 {chloride of merc.
60 chloride of sodium		{ oxygen 8	
		40	
		36	
		24	{ 72 {sulphate of soda.
308		308	
		308	

The last method of preparing chloride of mercury, which we shall describe, is by precipitation from the nitrate, (nitrate of the protoxyde). For this purpose, a solution of the nitrate of mercury, prepared with the usual precautions, must be added to a solution of chloride of sodium in 50 or 60 parts of water; chloride of mercury will be immediately precipitated, and nitrate of soda remain in the liquid. The quantity of chloride of sodium, recommended to be dissolved in the water, is equal in weight to the mercury employed in the preparation of the nitrate; a much smaller quantity would be sufficient, if the nitrate could be easily procured free from pernitrate, 262 parts of the nitrate reacting on 60 of the chloride; a small quantity of pernitrate of mercury is always formed, however, if the solution be not kept for a long time in contact with an excess of mercury, and every equivalent of the pernitrate reacts on 120 of the chloride, producing bi-chloride of mercury instead of calomel. The following diagram represents the reaction that takes place, between the nitrate of mercury and the chloride of sodium.

Before decomposition.		After decomposition.	
60 chloride of sodium	{ sodium chlorine nitric ac. 208 protox. of mercury	24	{ 86 {nitrate of soda.
262 nitrate of merc.		36	
		54	
		8	
	{ oxyg. 8		{ 236 {chloride of merc.
	{ merc. 200		
322		322	322

The chloride is precipitated in the form of a white powder, and must be washed and filtered with pure water.

The following diagram gives a view of the reaction that would take place between the mercurial salt and the chloride, supposing the whole of the mercury to be converted into pernitrate, (composed of two equivalents of nitric acid and one of the peroxyde,) and a sufficient quantity of chloride of sodium to be in solution.



324 pernitrate of mercury	{	mercury	200	}	bi-chloride of mercury.
		oxygen	8		
		nitric ac.	54		
		oxygen	8		
120 chloride of sodium= $60 \times 2$	{	nitric ac.	54	}	86 nitrate of soda.
		chlorine	36		
		chlorine	36		
		sodium	24		
		sodium	24		86 nitrate of soda.
444			444		444

The student will now perceive the great importance of attending to the quantity of oxygen and acid combined with the metal; as in the present instance, nothing would be obtained but bi-chloride of mercury and nitrate of soda, in place of chloride of mercury and nitrate of soda, if the mercury were converted into a pernitrate, instead of a nitrate of the protoxyde; and, moreover, no precipitate would take place, bi-chloride of mercury being soluble in water. He will likewise remark, that though chloride of mercury is prepared from a mixture of chloride of sodium and sulphate of mercury, bi-chloride of mercury is formed when persulphate of mercury is used.

Calomel is a semi-transparent mass, composed of short prismatic crystals. It is inodorous, insipid, and of an ivory color which deepens by exposure to the light. It is usually called an insoluble substance, and yet 1152 parts of water, at the boiling point, dissolve one of calomel, agreeably to Dr Paris. Mr Brande, however, affirms, that boiling water does not take up any portion of calomel, as he has shown by the use of sulphuretted hydrogen gas. This test produced no change of color in the water.

The specific gravity of calomel is 7.2. It has been imagined, by some, that a light buff color in calomel, denotes the absence of corrosive sublimate, but this is not a sufficient criterion; nor does it follow, that the snow-white color of some portions of this article, denotes the presence of corrosive sublimate. When the surface of massive, sublimed calomel is scratched, it always exhibits a buffy color.

The incompatible substances of this medicinal compound, are as follow: *alkalis* and *lime water* decompose it and turn it black, in consequence of the precipitation of the black oxyde of mercury; *soaps* and *sulphurets of potash and antimony* decompose it; *iron*, *lead*, and *copper* are also incompatibles, and, therefore, mortars of these metals should not be employed in preparing

compounds, of which calomel makes a part. It does not follow, however, that medicines should be prohibited, because chemically incompatible with this substance; for it is not improbable, that it undergoes decomposition in its passage through the alimentary canal, even when it answers our expectations most completely.

The *black wash* so well known to physicians, is formed by the addition of calomel to lime water, and results from the precipitation of the protoxyde of mercury.

*Sulphurets of Mercury.* Of these, we have two, viz. the *sulphuret* or *protosulphuret* and the *bi* or *deutosulphuret*. They are also called, the *black* and *red sulphurets*.

The *sulphuret* may be obtained, by passing a stream of sulphuretted hydrogen gas through a diluted solution of the protonitrate of mercury, or through water in which chloride of mercury, (calomel,) reduced to a fine powder, is suspended, by frequent stirring. The following diagram explains the reaction that occurs, when calomel is employed.

Before decomposition.			After decomposition.	
17 or 1 equiv. of	{hydrogen	1	37	muriatic acid.
sulph. hyd.	{sulphur	16		
236 or 1 equiv. of	{chlorine	36		
chloride merc.	{mercury	200	216	sulph. merc.
—		—	—	—
253		253	253	
—		—	—	—

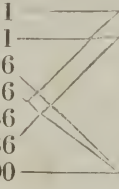
The sulphuret is thrown down in the form of a black powder, and muriatic acid remains in solution. If the sulphuret be separated by means of a filter, and exposed to heat, a portion of mercury will be obtained, and a bi-sulphuret will be formed. Thus, suppose two equivalents of sulphuret to be formed, equal to 432 parts; if one equivalent of mercury, 200, be detached, by the agency of heat, there will be left one equivalent of mercury, 200, and two equivalents of sulphur, 32, constituting one equivalent of bi-sulphuret of mercury, 232.

Brande affirms, that the black sulphuret of mercury, of the different colleges, is a mixture of sulphur and bi-sulphuret of mercury; but according to Paris, it is strictly a sulphuret, composed of one equivalent of each constituent. It was formerly called *æthiops mineral*, and was frequently employed as an alterative and anthelmintic.\*

\* The practice of using flowers of sulphur, to arrest or mitigate the violence of a mercurial salivation, is of ancient date; but I have never met with a satisfactory rationale of its *modus operandi*. The effect cannot

This compound may be also made, by rubbing together equal weights of mercury and sulphur, till the globules of mercury disappear. It is entirely soluble in a solution of pure potash, from which the acids precipitate it unchanged. It is insoluble in nitric acid.

The *bi-sulphuret of mercury* may be formed, by several methods. The easiest mode, perhaps, is to pass a stream of sulphuretted hydrogen gas through a solution of the nitrate of the peroxyde of mercury, or of the bi-chloride, as long as any precipitation takes place. The black matter, thrown down, is to be collected on a filter, and sublimed in a glass tube or any other convenient apparatus. In this operation it will assume a crystalline form, and a purplish-red color, which will be more striking if a portion of it is reduced to powder on a white ground. The annexed diagram is explanatory of the decomposition that occurs, when the bi-chloride is employed.

34 or 2 equiv. of sulph. hyd.	{	hydrogen	1		37 muriatic acid.
		hydrogen	1		37 muriatic acid.
272 or 1 equiv. of bi-chloride of mercury	{	sulphur	16		232 bi-sulph. merc.
		sulphur	16		
		chlorine	36		
		chlorine	36		
		mercury	200		
<hr/> 306			<hr/> 306		<hr/> 306

The whole of the mercury is precipitated in this process, and if the solution contain a considerable quantity of the bi-chloride, the muriatic acid that remains in the liquid will impart a deep red color to the blue infusion of cabbage, and decompose carbonate of lime rapidly, with evolution of carbonic acid gas.

The principal ore of mercury, and that which yields a large quantity of this metal, is a bi-sulphuret, and is known by the name of *native cinnabar*. The *vermillion* of the shops differs from this, only in its color; the cinnabar being in mass, has a reddish-brown color, but when it is reduced to a fine powder, it acquires a bright red color, and is then called vermillion.

The common method of preparing bi-sulphuret of mercury, or *artificial cinnabar*, is by melting 40 parts of sulphur in an iron

result from mere catharsis, or any purgative medicine would answer the purpose. Its action is no doubt chemical, and a sulphuret is formed, which, as such, has no salivant power. It does not, therefore, control the salivation actually existing; but, by combining with mercurial particles diffused through the alimentary canal, it removes the cause, and thus prevents the effect that would otherwise be continued for a much longer period.

cup over a chauffer, and adding 200 of metallic mercury, stirring constantly with an iron rod, till the mixture has assumed a uniform appearance, and taking care to apply only a moderate heat, to prevent it from burning; the mixture is extremely apt to take fire when the heat is too great, and the iron cup must be covered when this occurs, removing the chauffer for a short time. It must afterwards be reduced to powder, and sublimed in a close vessel. Eight parts more of sulphur are recommended to be taken, than is absolutely necessary to convert the metallic mercury into bi-sulphuret, to make up for a portion of sulphur which is always lost. The bi-sulphuret is obtained of a finer color when there is an excess of mercury, as a portion of this is sublimed along with the bi-sulphuret; the sublimation should be conducted slowly, the bi-sulphuret condensing in a crystalline cake, having a radiated appearance.

Bi-sulphuret of mercury is sometimes adulterated with red lead or chalk, either of which may be easily detected by exposure to heat on an iron plate; the pure bi-sulphuret is volatilised, while the chalk or oxyde of lead remains, the latter losing a portion of its oxygen and being converted into yellow oxyde of lead, if the temperature has been sufficiently high.

This bi or red sulphuret of mercury is inodorous, insipid, insoluble in water, alcohol, acids, and alkalis, although these last bodies decompose it when melted with it. It is also decomposed by nitromuriatic acid, which unites with the metal and disengages the sulphur. It is now employed only for the purposes of fumigation, in order thus to excite pyalism. For this end, a half drachm or more may be thrown on red hot iron. The mercury is volatilised, and thus the effects are produced.

*Carbonate of mercury* may be obtained, by adding a solution of the nitrate of mercury (protonitrate) to a solution of carbonate of potash, as long as any precipitation takes place. A double decomposition ensues, in which nitrate of potash and carbonate of mercury are produced; the former remains in solution, while the latter is precipitated. It consists of 22 parts carbonic acid, and 208 protoxyde of mercury, or one equivalent of each, as the following diagram evinces.

Before decomposition.		After decomposition.	
70 or 1 equiv. carbonate of potash	{ potash 48	102 nit. potash.	
262 or 1 equiv. nitrate of mercury	{ carbonic ac. 22		
	{ nitric acid 51	230 carb. merc.	
	{ protox. mer. 208		
332	332	332	

*Percarbonate of mercury* is formed, when a solution of the carbonate of potash is added to a solution of the pernitrate or bi-chloride of mercury. We could illustrate the reaction by a single diagram, but it might appear somewhat complicated and thus occasion embarrassment. We therefore give two diagrams, in the first of which, the changes resulting from the mutual action of the water (of the solutions) and the bi-chloride, are plainly seen. The second shows the decomposition produced by adding to the results of the operation exhibited in the first diagram, a given quantity of carbonate of potash.

Before decomposition.		After decomposition.	
18 or 2 equiv. of water.	{ hydrogen 1	1	37 muriatic acid.
	{ hydrogen 1	1	37 muriatic acid.
272 or 1 equiv. bi-chloride of mercury	{ oxygen 8	8	216 peroxyde of merc.
	{ oxygen 8	8	
	{ chlorine 36	36	
	{ chlorine 36	36	
	{ mercury 200	200	
<hr/> 290 <hr/>	<hr/> 290 <hr/>	<hr/> 290 <hr/>	

The muriatic acid and peroxyde of mercury, thus produced, react on two equivalents of carbonate of potash, giving rise to muriate of potash, and *per* or *bi-carbonate* of mercury.

Before decomposition.		After decomposition.	
140 or 2 equiv. carb. potash	{ potash 48	48	85 mur. potash.
	{ potash 48	48	85 mur. potash.
74 or 2 equiv. mur. acid	{ carb. ac. 22	22	260 percarb. mercury.
	{ carb. ac. 22	22	
216 or 1 equiv. perox. mer.	{ mur. acid 37	37	
	{ mur. acid 37	37	
<hr/> 430 <hr/>	<hr/> 430 <hr/>	<hr/> 430 <hr/>	

Expose a small quantity of any of the precipitated carbonates of mercury to heat, in a tube over a spirit lamp; carbonic acid and oxygen gases will be disengaged, and small globules of metallic mercury soon appear a little above the part of the tube that is held over the flame.

*Acetate of mercury* is easily prepared, according to the process of the Edinburgh college. Three parts of mercury are to be dissolved in four and a half of their diluted nitrous acid, (composed of equal weights of water and the strong fuming acid,) and



the solution must be added to an equal weight of the acetate of potash, dissolved in 32 times its weight of water; nitrate of potash remains in solution, and acetate of mercury is deposited in small crystals as the liquid cools.

Before decomposition.		After decomposition.	
99 or 1 equiv. of	{ potash	48	102 nit. potash.
acetate of pot.	{ acetic acid	51	
262 or 1 equiv. of	{ nitric acid	54	
nit. mercury	{ oxyde of mer.	208	259 acet. merc.
361		361	361

In preparing the nitrate, the usual precautions must be taken to prevent the formation of peroxyde of mercury.

*Peracetate of mercury* may be obtained, by digesting the peroxyde of mercury in acetic acid, till it will not dissolve any more. The acetate of mercury constituted the basis of *Keyser's pills*, a celebrated remedy for syphilitic affections. It is composed of one equivalent of acetic acid, 51, and one of protoxyde of mercury, 208.

The method of preparing *bi-cyanide of mercury*, has been already described.

*Cyanate of mercury*, or *fulminating mercury*, as it is usually termed, is prepared by mixing perntrate of mercury, dissolved in an excess of acid, with alcohol. For this purpose, 100 grains of mercury may be digested with an ounce and a half, by measure, of strong nitric acid in a Florence flask, till they are dissolved, and the solution after it has been allowed to cool a little, added in small quantities at a time, to two ounces of alcohol in another flask, exposing it afterwards to a very gentle heat over a chauffer, till white fumes begin to appear. If the action should proceed very violently, it must be moderated by adding a little alcohol; and if it do not commence soon after the mixture is exposed to heat, a few drops of strong nitric acid poured in by a pipette, will soon cause the white fumes to appear. These fumes are extremely heavy, and may be poured from one vessel to another, or through a funnel; their composition has not been ascertained. When the reaction has ceased, a quantity of a white crystalline powder is found to have been deposited, which is the fulminating mercury; it must be washed on a filter with water, and allowed to dry by exposure to the open air. It should never be put into a bottle with a ground stopper, as it detonates by friction and percussion, but should be merely folded in a piece of paper, and kept in a wide-mouthed

vial with a cork fitting loosely to it. The cyanic acid in this compound, is formed by the decomposition of part of the nitric acid and the alcohol, the nitrogen being derived from the acid, and the carbon from the alcohol. It was formerly called *fulminic acid*, simply from the fulminating property of the powder.

Place ten or twelve grains of fulminating mercury on a block of iron, and touch it with a red hot cinder; it will immediately  
 Exp. detonate with a quick but not a sharp or loud report, and a bluish-white flame, its elements being separated from one another. It may be detonated also by striking it with a hammer, and is the basis of the detonating mixture employed for the anti-corrosive percussion caps.

If two or three grains be put into a dry Florence flask, to which a brass cap and stop-cock have been fitted, and exposed to heat over a spirit lamp, after the air has been exhausted as  
 Exp. completely as possible, a flash of light will soon be perceived, but no report heard, and metallic mercury will be deposited on the inside of the flask. If the air of the apparatus be not exhausted, a loud explosion will take place, and the flask be blown to pieces. Seven grains of fulminating mercury were found sufficient to produce the same effect when the flask had been exhausted of air as completely as possible.

An attempt was made to substitute this article for common gunpowder, but its tremendous force defeated the enterprise; several fatal accidents occurred, however, before the design was abandoned.

*Phosphate of Mercury.* This salt was formerly called *phosphorated mercury*, and once stood high as an anti-venereal medicine. It is composed of phosphoric acid, one equivalent, or 28 parts, and protoxyde of mercury, one equivalent, or 208 parts; and may be made by the double decomposition of phosphate of soda and protonitrate of mercury, mixed in solution in the proportion of one equivalent of each article. The results are nitrate of soda and phosphate of mercury.

*Muriate of ammonia and mercury,* the *hydrargyrum precipitatum album* of the London college, and the *sal alembroth* of the older chemists, may be prepared by adding a solution of carbonate of potash, to a solution of the bi-chloride of mercury in muriate of ammonia, as long as any precipitation takes place. The following are the proportions recommended: four ounces of the muriate of ammonia are to be dissolved in four or five pints of distilled water, and six ounces of the bi-chloride of mercury must then be dissolved in this solution, separating the white precipitate that is thrown down, on adding the solution of carbonate of potash, by filtration, and washing it with water on the filter.

In this process, the bi-chloride may be considered to have been converted into muriatic acid and peroxyde of mercury, by reacting on two equivalents of water, and the potash of the carbonate uniting with the acid, the peroxyde is precipitated in combination with the muriate of ammonia. This preparation is employed, only as an external application, in the form of ointment.

*Iodide of mercury* may be obtained, by adding a solution of the hydriodate of potash to a solution of the nitrate of mercury (protonitrate); it is immediately precipitated in the form of a yellow powder. The potash combines with the nitric acid of the nitrate, the hydriodic acid and oxyde of mercury mutually decomposing each other, and forming water and iodide of mercury.

*Periodide of mercury* is formed and precipitated of a red color, when a solution of pernitrates of mercury is added to a solution of hydriodate of potash, two equivalents of the latter being decomposed by one of the pernitrates, and the usual reaction taking place between the acid and the oxyde.

It may also be obtained, very easily, by heating iodine with rather more than twice its weight of mercury, in a glass tube; a brisk ebullition takes place, and periodide of mercury is sublimed, condensing on the sides of the tube, and assuming a very rich crimson color, as it cools. An excess of mercury should be used, as the periodide is not then liable to be mixed with any uncombined iodine, which would prevent its bright color from being so distinctly seen.

---

## CHAPTER II.

### Of Silver.

THIS, in common with mercury, gold, and platina, was formerly distinguished by the appellation, *precious metal*. But if either of these be compared with iron, in all the points of real usefulness, we shall be disposed to set upon the latter, a much higher value than upon either of the former.

Silver frequently occurs in the metallic state, in silver mines, and it is also found in combination with other metals, and with sulphur.

Pure silver is always white, has a specific gravity of 10.5, and its equivalent is 110. It requires a heat of from  $20^{\circ}$  to  $22^{\circ}$  of

Wedgewood to fuse it, and at a still higher temperature, it is converted into vapor. It is susceptible of a very high polish, has considerable tenacity, and in respect of malleability and ductility, is inferior only to gold. When pure, it is very soft, and may be cut with a knife. It does not rust on exposure to air and moisture, nor is it oxydated by fusion in open vessels; hence its fitness for crucibles and other utensils that are subjected to a strong heat. In the form of leaves or fine wire, it may be burnt by means of common electricity, galvanism, or the oxy-hydrogen blowpipe; when thus burnt, it gives out vivid scintillations of a greenish-white color.

Silver is not acted upon by sulphuric acid, unless heat be employed. Nitric acid dissolves it, but the action is more rapid, if a spirit lamp be used in the process.

Pure silver is usually obtained, on the small scale, by precipitation from a solution of the nitrate, introducing a piece of metallic copper for this purpose, which unites with the oxygen and acid of the nitrate, while the silver is precipitated in small crystalline grains, which must be washed on a filter with distilled water.

It may be precipitated in the metallic form by mercury also, exposing the silver that is obtained in this manner to heat, in a crucible till it is melted, that the mercury which is mixed with it may be expelled. This process is seldom resorted to, however, except for the purpose of showing the arborescent form in which the silver is separated by the mercury. It is in this manner that the *arbor dianæ*, as it has been termed, is usually prepared; 50 grains of the fused nitrate of silver dissolved in two or two and a half ounces of water, and put into a glass with 100 grains of mercury, do very well for this purpose. The crystals of silver are deposited above the mercury and rise slowly in the liquid, part of the latter metal being slowly dissolved, as it combines with the oxygen and acid previously in combination with the silver.

Silver is frequently alloyed with a small quantity of copper, and we now state the processes by which they may be separated. The one usually followed, consists in exposing the alloy to the action of the air, at a high temperature, along with several times its weight of lead, the copper and the lead being oxydated and fused, while metallic silver remains. For this purpose, the alloy must be placed on a cupel, a small porous crucible made of bone ashes, and exposed to heat in the muffle of a cupellation furnace; or, some bone ashes may be spread on a tile or flat earthen dish, if a cupel cannot be procured. The quantity of lead required, depends on the richness of the alloy;



if it contain about a tenth part of copper, which is usually the case, six or seven times its weight of lead will be found quite sufficient; but if more copper be present, it will be necessary to use a larger quantity of lead.

The use of the lead in this process, is to form a fusible compound with copper, during the oxydation of both metals, which is speedily absorbed by the bone ashes, while the pure silver remains above. If too small a quantity of lead be employed, a brown crust gathers on its surface which is not fused, and prevents the farther oxydation of the rest of the lead and copper. The usual precautions must be taken in bringing the muffle to a proper temperature, and the process is known to be completed, by the fulguration or brightening, as it is termed, which is seen when the last portions of lead and copper are oxydated and removed.

On taking out the cupel, the silver is seen in the form of a metallic globule, of a rich, white color and great lustre; the fused oxydes give the cupel a very dark appearance, where they have been absorbed, deeper in proportion to the quantity of copper which the alloy may have contained. The experiment may be made with a few grains of the alloy, or a much larger quantity may be employed.

Another process, for separating silver from copper, may be adopted where there is no cupellation furnace. Digest the alloy with three parts of nitric acid, diluted with twice its bulk of water, in a glass flask or evaporating bason, till it is dissolved, and pour the liquid into a dilute solution of chloride of sodium, containing a quantity of the chloride equal to about twice the weight of the alloy. The nitrates of copper and silver react upon the chloride of sodium, nitrate of soda and chloride or muriate of copper being formed and remaining in solution, while chloride of silver is precipitated. The diagram represents the reaction that takes place between the nitrate of silver and chloride of sodium.

Before decomposition.			After decomposition.	
60 or 1 equivalent chlor. sodium	{	sodium	24	86 nitrate of soda.
		chlorine	36	
		nitric acid	54	146 chloride silver.
		oxygen	8	
		silver	110	
172 or 1 equiv. nit. of silver				
232			232	232

The chloride of silver is then to be washed on a filter with water, till the liquid that passes through, gives no traces of cop-



per; mix it with its own weight of dry carbonate of potash, and put the whole into a bottle or any kind of a glass vessel, which being placed in a crucible and surrounded with sand, must be exposed to a strong heat in a furnace. The potassium of part of the carbonate, unites with the chlorine of the chloride, forming chloride of potassium, which fuses along with the rest of the carbonate, the glass of the bottle or tube containing the mixture and part of the sand, forming a well defined ball of glass; and on breaking into it, the metallic silver will be found, generally, in a solid mass at the lower part of the ball.

It is generally stated, that there is a considerable portion of the silver lost, in reducing the chloride by carbonate of potash; but almost the whole of it is obtained by conducting the process in the manner here described. The following diagram represents the reaction that takes place, between the potash of the carbonate and the chloride of silver.

Before decomposition.			After decomposition.	
48 or 1 equiv. pot.	{ oxygen	8	8 oxy. disengag.	
	{ potassium	40		
146 or 1 equiv. of	{ chlorine	36	76 chlor. potassi.	
chloride of silver	{ silver	110	110 silver.	
194		194	194	

To prepare metallic silver from the chloride, Dr Ure recommends 100 of it to be fused in a crucible, with 19.8 of pure lime and 4.2 of charcoal.

Mr Keir discovered, that an acid liquor, composed of eight parts of sulphuric acid and one of nitre, has the property of dissolving silver, while it does not act upon copper, and has accordingly recommended it to be employed for removing silver from plated goods. The action of the liquid should be assisted by a moderate heat, not exceeding that of boiling water, and the silver precipitated by a solution of common salt, reducing the chloride that is precipitated in the usual manner.

There is but one *oxyde of silver*, and that is composed of one equivalent of oxygen, 8, and one of silver, 110. It is of a brownish color, and may be obtained, by adding a solution of the nitrate of silver, to a solution of barytes; the oxyde of silver is immediately precipitated, while nitrate of barytes remains in solution. The alkalis and alkaline earths also precipitate oxyde of silver from the nitrate.

Fulminating silver is a compound of ammonia and oxyde of silver, usually prepared by pouring water of ammonia on oxyde of silver, precipitated from the nitrate by lime water, after

washing it on a filter. In twelve hours the liquid must be cautiously decanted, and the black, fulminating compound that remains, allowed to dry spontaneously on small pieces of filtering paper.

As many accidents have occurred during the preparation of this compound, even with those who have been accustomed to chemical manipulations, it will be better for the beginner to pass over this process. It detonates, when touched with any hard body, and if a few grains be used, the explosion is extremely violent. The liquid, also, if gently heated, affords a still more dangerous compound, which explodes even when touched under the surface of the liquid\*

The most delicate test of silver, in solution, is muriatic acid, or a solution of any muriate, chloride of silver being immediately thrown down in the form of a white, curdy precipitate, which becomes of a dark purple color, on exposure to the air. Sulphuretted hydrogen and solutions of the hydrosulphurets, give a black precipitate of sulphuret of silver; arsenite of potash causes a yellow precipitate of arsenite of silver. Small metallic globules of silver, obtained by decomposing any of its compounds on charcoal, before the blowpipe, are easily recognised by their brilliant white color, their hardness, and the manner in which they rest above the surface of the charcoal, not being embedded in it like many other metals reduced on charcoal, but rising above it, and adhering only by a small point.

The most important salt of this metal is the *nitrate of silver*. It is prepared, by digesting silver in a glass vessel with one and a half times its weight of nitric acid, diluted with an equal bulk of water, and evaporating the solution to dryness. One portion of the acid affords oxygen to the silver, and nitric oxyde gas is disengaged, the oxyde formed in this manner combining with the acid that is not decomposed. When dissolved in less than its weight of hot water, the solution affords tabular crystals as it cools.

The fused nitrate of silver of the different colleges, or *lunar caustic*, is prepared by melting the crystallised nitrate in a porcelain crucible, capable of containing four or five times as much as is employed, heating it very gently at first, to prevent it from boiling over, and pouring it, whenever it becomes quite liquid, into cylindrical moulds well greased with a little tallow. The operation must be conducted with care, to avoid the very

\* A very small quantity should be kept in a bottle or paper, at a time; for even the agitation of a vial causes its explosion. Professor Hare was very much injured, a few months ago, by an explosion of this sort.

caustic sparks that are occasionally thrown out of the crucible, during the fusion of the nitrate.

The freshly prepared lunar caustic is of a grayish-white color, but becomes darker, by exposure to the light. Its taste is intensely bitter, austere and metallic, and owing to the presence of copper, it is usually deliquescent. It is composed of 51 nitric acid, one equivalent, and 118 oxyde of silver, one equivalent. It is soluble in its own weight of water at 60°, and is also soluble in alcohol.

Nitrate of silver stains the skin and hair black, and is decomposed by sulphuric and muriatic acids, solutions of the alkalis and earths, sulphuretted hydrogen, hydrosulphurets, by many of the metals, and a great number of astringent vegetable solutions.

These substances may therefore be regarded as incompatible with this salt. It is used chiefly as an escharotic, though occasionally employed internally, in very small doses. When administered for a considerable length of time, it has caused purple tinges in the skin, owing to a deposit of the oxyde of silver. In many cases in which it is used, free muriatic acid is known to exist in the stomach; of course, the salt must then be decomposed and a muriate be formed.

Common *marking ink* is composed of a solution of this salt, thickened with a little mucilage, and the preparatory liquid with which the part to be marked is previously moistened, is a solution of carbonate of soda, also thickened with a little mucilage. 100 grains of the fused nitrate may be dissolved, for this purpose, in distilled water, and two or three drachms of mucilage added to the solution, keeping it in a bottle; for the preparatory liquid, half an ounce of the carbonate of soda may be dissolved in two or three ounces of water, adding half an ounce of mucilage to the solution.\*

Nitrate of silver is sometimes adulterated with copper, in which case it is deliquescent, and its solution gives a blue color, when saturated with ammonia. When this salt has been taken in an overdose, a solution of common house salt, (muriate of soda,) is the best antidote. So entirely does this agent decompose the nitrate and destroy its deleterious properties, that if a solution of the lunar caustic be filtered through common salt, the liquor that passes the filter, may be drank with impunity. Thenard, in the course of one of his lectures, drank the contents

\* In using this *marking* or *indelible ink*, the linen is first impregnated with the solution of carbonate of soda and dried; then the writing is to be executed with a fine stiff pen, and on drying it, the letters will be perfectly legible.

of a tumbler, supposing it to be water; it proved to be a solution of nitrate of silver, and he immediately swallowed another tumbler full of a solution of muriate of soda; no bad consequences followed. On the same principle, a solution of nitrate of silver is employed, to detect the presence of muriatic acid, either free or combined, in mineral waters, and vice versa.

*Sulphate of silver* may be obtained, by digesting metallic silver in sulphuric acid, one portion of the acid being decomposed, affording oxygen to the metallic silver, while sulphurous acid gas is disengaged, and the rest of the acid combines with the oxyde. It may be prepared, also, by adding a solution of the sulphate of soda to a solution of the nitrate of silver; nitrate of soda remains in solution and sulphate of silver is precipitated. It is employed as a test of muriatic acid.

*Phosphate of silver* is precipitated, when a solution of the phosphate of soda is added to a solution of the nitrate of silver, a double decomposition taking place, and nitrate of soda remaining in solution.

*Carbonate of silver* may be obtained, by adding a solution of an alkaline carbonate to a solution of the nitrate of silver.

*Cyanate of silver*, another detonating compound of silver, less dangerous than the one already described, may be prepared by dissolving metallic silver in strong nitric acid, and adding the solution to alcohol, proceeding in the manner directed for the preparation of cyanate of mercury and using the same proportions of metal, acid, and alcohol. It explodes much more violently than cyanate of mercury, and should be handled with still more precaution. It should not be touched with any thing but a piece of paper or card, except for the purpose of experiment. It is detonated by heat, friction, galvanism, percussion, and several of the acids, producing a very sharp report. It is a cyanate of the oxyde of silver, and contains one equivalent of cyanic acid, 34, and one of oxyde, 118.

*Chloride of silver*, already noticed, is always formed, when muriatic acid or a solution of any muriate is added to a solution of a salt of silver; it is precipitated in the form of a white, curdy looking powder. Exposed to the direct rays of the sun, it soon takes on a purple or black color, owing to the decomposition of part of the chloride. It is easily fused by a temperature of about  $500^{\circ}$ , and a mass like horn is formed, as it cools, which has been called *luna cornea*, or *horn silver*.

*Iodide of silver* is formed, when the hydriodate of potash is mixed with a solution of the nitrate of silver. The hydriodic acid of the hydriodate, is decomposed, its hydrogen joins the oxygen of the oxyde of silver, while the silver, thus liberated,



unites to the iodine of the decomposed hydriodic acid, and iodide of silver is the product. It is of a greenish-yellow color, insoluble in water and ammonia, and contains one equivalent of each of its elements.

Silver also combines with cyanogen and sulphur, forming cyanide and sulphuret of silver.

---

## CHAPTER III.

### Of Gold.

THIS term is of Saxon origin, and was employed, originally, to denote the purest, heaviest and most valuable of the metallic bodies.\* It has been found only in the metallic state, either pure or combined with other metals. It occurs in a variety of forms, as in fine grains, crystals, in large masses, &c. It is found not only in Africa, Hungary, and other foreign countries, but also in several of these United States.

Gold is the only metal that has a perfectly yellow color, and by this it is distinguished from all other simple metallic bodies. It is susceptible of a high polish, but its brilliancy is less than that of steel, silver, and mercury. It is more ductile and malleable than any other metal,† but less tenacious than some of the metals. When pure, it is very soft and flexible, fuses at  $1200^{\circ}$ , assuming at the same time a bluish-green color. Its specific gravity is 19.3, and its equivalent 200. It may be exposed to air and moisture for ages, without undergoing any change. It is not acted on by any single acid, however strong it may be, and its proper solvent appears to be chlorine.

When gold is alloyed with copper, it may be separated by the process of cupellation, proceeding in the manner that has been described for the cupellation of silver; 15 or 20 grains of an alloy, containing about a tenth of its weight of copper, will be sufficient to show the nature of the process on the small scale.

Silver and platina cannot be removed in this manner, as neither of these metals is oxydated by exposure to a high tem-

\* Platina, since discovered, is a heavier metal.

† So great is its malleability, that 280,000 leaves of gold piled upon each other, measure but an inch in thickness, and the gilding on silver wire, has only one-twelfth of the thickness of gold leaf.



perature. Platina is not often alloyed with gold, and silver may be separated by the operations of *quartation* and *parting*. Quartation consists in fusing the alloy with three times its weight of silver, by which the particles of the gold are separated to a greater distance from each other, and prevented from covering or protecting any of the particles of silver from the action of nitric acid. Parting consists in boiling the alloy in seven or eight times its weight of nitric acid, to remove the silver, diluting the acid with an equal bulk, or rather more, of water, and repeating the operation with a smaller quantity of acid, till the whole of the silver has been extracted. As gold is not soluble in nitric acid, all that the alloy may have contained, is left in a porous mass of the same form as the original alloy, or reduced to powder.

When gold is exposed to an intense heat by the oxy-hydrogen blowpipe, it is dissipated in vapor, and a purple powder may be collected, which has been regarded as an oxyde of gold. Its composition is still uncertain.

*Oxydes of Gold.* Agreeably to Berzelius, there are three oxydes of gold, viz. the protoxyde, deutoxyde, and peroxyde, but there is much obscurity respecting their actual composition. That which is best understood, is the peroxyde, which is also called *auric acid*, on the supposition that it can combine with bases, forming a class of salts, named *aurates*.

*Peroxyde of gold* may be obtained, by adding a solution of pure potash to a solution of the chloride of gold, taking care not to add excess of alkali; muriate of potash remains in solution, and the peroxyde is precipitated in combination with a portion of water, the usual reaction taking place between the metallic chloride and part of the water that is decomposed.

*Chloride of gold* may be obtained, by digesting small fragments of gold in a mixture of one part of nitric acid and two of muriatic acid, evaporating the solution to dryness with a very gentle heat, to expel any excess of acid. The gold is dissolved by the chlorine, evolved by the mutual reaction of the nitric and muriatic acids. If the heat be too strong, the chlorine will be dissipated, and nothing will remain but metallic gold; great caution is, therefore, required, in evaporating the chloride to dryness.

A weak solution of the chloride of gold may be obtained, by shaking gold leaf with a solution of chlorine in water. None of the acids act upon gold, except the nitromuriatic and a mixture of chromic and muriatic acids, and in both cases, chlorine may be considered the actual solvent, part of the oxygen of the nitric or chromic acid combining with the hydrogen of a portion of muriatic acid, and disengaging this element.

Exp.

Digest the dry chloride of gold in water, and filter the liquid;  
 Exp. a solution of chloride of gold is obtained, of a deep reddish-brown color. If it is concentrated and then set aside to cool, small crystals of chloride of gold are deposited.

*Ammoniuret of gold*, or *fulminating gold*, may be obtained, by adding ammonia to a solution of the bi-chloride of gold; a portion of water being decomposed and peroxyde of gold precipitated in combination with ammonia, while muriate of ammonia remains in solution. It is of a reddish-brown color, and should be allowed to dry on a filter at natural temperatures, after washing it with water. It detonates violently by friction and percussion, or when exposed to heat, its elements being separated from one another; it must never be touched with any hard substance, except for the purpose of experiments. It should be kept in a wide-mouthed bottle, closed tightly with a cork, and in small packets of paper.

Add a solution of the green sulphate of iron (protosulphate)  
 Exp. to a solution of the chloride of gold; metallic gold is precipitated, water being decomposed and its hydrogen combining with the chlorine, while the oxygen converts the protoxyde of iron into peroxyde.

Digest the precipitate, thrown down in this manner, in diluted  
 Exp. sulphuric or muriatic acid, to remove any oxyde of iron that may have been thrown down along with the gold, and the latter will be obtained in a state of great purity.

Add a solution of the muriate of tin (protomuriate) to a solution of the chloride of gold; a purple precipitate is immediately  
 Exp. thrown down, which has been regarded as a compound of peroxyde of tin and the oxyde of gold. It is well known by the name of *purple of Cassius*, and is used for communicating a rich red or pink color to glass.

Touch a piece of crystallised borax with a glass rod that has  
 Exp. been dipped into a solution of the chloride of gold, and on exposing it to heat before the blowpipe, on a piece of earthen ware or baked clay, or fusing it in a crucible, a glass will be obtained which will have a very rich red or pink color.

*Ethereal solution of gold*, or *potable gold*, which is used for gilding a number of substances, may be easily prepared, by shaking a strong solution of the chloride with an equal bulk of pure ether, the ethereal solution collecting above, and a heavier liquid remaining below. It should be decanted immediately,  
 Exp. and kept in a stoppered bottle, enclosed in a case, to guard it from the action of light. If phosphorus or polished metals be dipped into this solution, they are instantly coated with gold.

The ethereal solution has been used in medical practice; so also have the peroxyde of gold and the choride of that metal. They have been employed as substitutes for mercury, in the treatment of venereal affections, but at present, they are seldom resorted to.

## CHAPTER IV.

### Of Platina.

THIS metal is found only in the metallic state, in union with other metals, as copper, iron, lead, gold, silver, &c. It is met with in Brazil, Peru, and South America, in the form of rounded or flattened grains, of a metallic lustre and white color, mixed with silicious matter. It is the heaviest of the metals, having a specific gravity of 21.5. Its equivalent is 96. Air and moisture exert no action upon it, and it is not fused, on exposing it to the strongest heat of a smith's forge; small portions may be easily melted, however, by drawing it into thin wires, and holding them in the flame of the oxy-hydrogen blowpipe. It may be welded like iron by exposing it to a high temperature, and it is in this manner that platina crucibles and other vessels made of this substance are formed. Platina is less malleable than gold or silver. It may be drawn into wires, the diameter of which does not exceed the two-thousandth part of an inch. It is a less perfect conductor of caloric, than most other metals.

Platina may be obtained from its ores, by digesting them in a mixture of three parts of muriatic and one of nitric acid, adding a solution of the muriate of ammonia to the solution of the chloride of platina obtained in this manner, and exposing the yellow precipitate, that is thrown down, to a red heat in a crucible, placing it in a chauffer on the open fire. It is regarded, as a compound of muriate of ammonia and perchloride of platina, and is completely decomposed by heating it; muriate of ammonia and chlorine being disengaged, while nothing remains but metallic platina, of a dull leaden color, and in the same minute state of division in which it is precipitated.

It is in this manner, that the *spongy platina*, which becomes incandescent on bringing it into contact with air and hydrogen gas, is prepared.

*Oxydes of Platina.* Berzelius has two oxydes of this metal, viz. the protoxyde and peroxyde. The former is prepared, by

the action of potash on the chloride of platina, in solution. The water of the solution is decomposed, its hydrogen changing the chlorine into muriatic acid, which combines with the potash, while its oxygen goes to the platina, and forms the protoxyde of that metal, composed of eight parts oxygen, and ninety-six platina. It is of a blackish color, and is reducible by a strong heat.

The *peroxyde of platina* is not easily obtained in a pure state, from its tendency to form triple salts when thrown down from any solution in which it may exist. It has been procured, by boiling the perchloride of platina with sulphuric acid, and decomposing the persulphate, formed in this manner, by a solution of potash. It contains two equivalents of oxygen to one of metal.

Platina may be detected in solution, by the dark, port wine color which solutions of hydriodic acid communicate to any liquids containing it. A solution of the muriate of tin renders the liquid of a bright red color. A solution of potash gives an orange-colored precipitate; but with soda, the liquid remains perfectly transparent. Sulphuretted hydrogen gas gives a black precipitate; and on evaporating any of the solutions of this metal to dryness, metallic platina may be obtained by exposing the residuum to a strong heat.

*Sulphate of platina* may be formed, by transmitting a stream of sulphuretted hydrogen gas through a solution of the chloride of platina, and digesting the precipitate in nitric acid. It is soluble in water, alcohol, and ether.

*Fulminating platina* may be obtained, by adding ammonia to a solution of the sulphate of platina, and boiling the precipitate that is thrown down, in a solution of potash; it must then be washed on a filter with water, and allowed to dry. It explodes violently when heated to the temperature of  $400^{\circ}$ , and is analogous in its constitution, to similar compounds of gold and silver.

*Chloride of platina* may be prepared, by digesting metallic platina in a mixture of nitric and muriatic acids, in the proportions recommended for separating it from its ores, evaporating the solution to dryness by a gentle heat, dissolving the residuum in water, and filtering the solution.

In this process, the chlorine set at liberty, by the reaction of the nitric and muriatic acids, dissolves the platina, none of the acids having any action upon this metal when perfectly pure, though it has been ascertained, that nitric acid can dissolve it, when an alloy, containing a small portion of this

metal, is digested in it. We noticed the application of this article, as a test of potash, when treating of the salts of potassium.

A singular compound has been described by Mr J. Davy, composed of platina, oxygen, nitrous acid, and carbon. It is prepared, by boiling sulphate of platina, (prepared as above described,) in strong alcohol, and drying the black powder that is precipitated. It is distinguished by becoming incandescent, and being completely decomposed, when it is put on a piece of bibulous paper moistened with alcohol, a hissing noise being produced at the same time, and nothing remaining but metallic platina.\*

## CHAPTER V.

### Of Nickel.

THIS metal is classed, by some authors, among those which require the aid of charcoal for their reduction. But as it is reducible by a very intense heat alone, we think proper to place it in connexion with those metallic bodies, which admit of separation from the oxygen of their oxydes, by the unassisted power of elevated temperature.

The equivalent of nickel is 26, its specific gravity 8.5. At a very high temperature, it may be fused, and then it absorbs oxygen. It is attracted by the magnet.

Nickel is a constituent of meteoric iron, and it occurs also, in the copper colored mineral of Westphalia, called *copper-nickel*.

This metal is usually obtained from the impure native sulphuret, the ore being heated to drive off the sulphur and arsenic, and the residuum intimately mixed, and exposed to a very high temperature, with twice its weight of black flux.

The nickel obtained in this manner, is still very impure; it may be separated from other metals with which it may still be combined, by the following process. Reduce a quantity of the alloy of arsenic and nickel, (well known in commerce by the name of *speiss*,) to powder, and boil it in diluted sulphuric acid, till it is dissolved, adding small quantities of nitric acid, from time to time, to promote the oxydation; a deep green colored liquid is obtained in this manner, containing sulphate of nickel

\* For some important facts, on the application of platina to the making of cooking and other utensils, the reader may consult the *Emporium of Arts and Sciences*, published at Philadelphia, in 1813.



in solution, which must be decanted from the arsenious acid that remains undissolved. After concentrating it by evaporation, and putting it aside to cool, green crystals of sulphate of nickel are deposited. These must be dissolved in water and carbonate of nickel precipitated from the solution, by carbonate of soda, and on exposing the light green colored powder which is precipitated, to a very strong heat, for one or two hours with charcoal in a crucible, a button of metallic nickel is obtained.

By evaporating the liquid that remains after the first crystallisation of sulphate of nickel, more crystals of this salt are procured; but when this has been repeated two or three times, the crystals that are deposited are composed of sulphate and arseniate of nickel. From these, pure sulphate of nickel may be obtained, by dissolving them in water, passing a current of sulphuretted hydrogen through the solution, as long as any precipitation takes place, filtering and evaporating the clear solution, concentrating the liquid that is obtained after dissolving the green residuum in water, and filtering it.

*Oxydes of Nickel.* Nickel is susceptible of two degrees of oxydation; the protoxyde contains one equivalent of oxygen, 8, and one of nickel, 26; while the peroxyde is estimated to be composed of the same proportion of metal, to one equivalent and a half of oxygen.

The protoxyde may be obtained, by heating the carbonate of nickel to redness, in an open crucible. Its color is gray, and the most of its salts have a green color.

The *peroxyde* may be prepared, by transmitting chlorine through water, in which the protoxyde is diffused in fine powder; a portion of this liquid is decomposed, the hydrogen combining with the chlorine and the oxygen with the oxyde of nickel.

The nitric and nitromuriatic acids are the best solvents of this metal; sulphuric acid has little action on it, but combines readily with the oxyde that is formed, on adding a little nitric acid.

---

## CHAPTER VI.

Of Palladium, Rhodium, Osmium, and Iridium.

THESE metals are found in the ores of platina, and have hitherto been obtained in very minute quantities. When the ore is digested in nitromuriatic acid, the platina, together with pal-

ladium, rhodium, and other metals present, is dissolved, while a black powder is left, which consists of osmium and iridium.

As these metals cannot be exhibited, and as they do not furnish any article of interest to medicine or practical chemistry, I will not detain the reader with the details on these subjects, that may be found in larger treatises.

## ALLOYS. AMALGAMS.

*Alloys* are usually prepared, by melting together the different metals of which they are composed; and those in which mercury forms a constituent part, which are termed *amalgams*, may be formed, by dissolving the different metals which they may contain in mercury, when a large proportion of this metal is required; and by adding it to them, if the quantity be comparatively small.

In the following paragraphs, the composition of some of the most important alloys is stated, small quantities of which may be easily prepared in the manner we have described.

*Brass* is composed of copper and zinc, usually in the proportion of three to one, and occasionally contains a small amount of other metals. A small quantity may be easily prepared, by melting two parts of copper and one of zinc in a crucible, covering the mixture with a little salt and charcoal, to protect it from the action of the air, and pouring out the alloy as soon as the metals have been completely melted, and stirred together with an iron rod. It is better to use an excess of zinc, as a considerable quantity is always volatilised by the heat.

On the large scale, brass is prepared in a different manner, as there would be a great loss of zinc, if the pure metals were melted together in a crucible. The impure oxyde of zinc is mixed with charcoal, and exposed to heat in a covered crucible, pieces of sheet copper being laid over the mixture. The zinc is reduced and converted into vapor, combining with the copper as it is disengaged, and forming brass; when the process is conducted in this manner, the brass is said to be prepared by *cementation*.

*Dutch gold* and *pinchbeck* are composed of the same metals, and contain a large quantity of zinc.

*Bronze* is an alloy of copper and tin, and is used for making statues, cannons, bells, and for a variety of other purposes. The proportions usually employed, are nine parts of copper to one of tin.

*Speculum metal*, (the alloy used for making the specula of reflecting telescopes,) is composed of two parts of copper and one of tin.

a small quantity of zinc and arsenic, about one part of each to fifty of the alloy, may be added, to improve its lustre.

*Type metal* is, usually, an alloy of lead and antimony, the latter being employed to communicate hardness; and occasionally, copper and brass are added, in small proportions. For making the small letter, more antimony is used than for the large, in order to increase the hardness.

*Pewter* is composed of about twenty parts of tin and one of copper; small quantities of antimony and bismuth are occasionally melted with it, and give it a whiter color. Inferior kinds of pewter are made, with about a fifth of its weight of lead.

*Plumber's solder* is composed of equal parts of tin and lead.

The *fusible metal* is an alloy of eight parts of bismuth, five of lead, and three of tin, melted together in a crucible. It fuses at a temperature of  $210^{\circ}$ , and teaspoons made of it, speedily melt in boiling water.

*Gold coin*, for the most part, is an alloy of eleven parts of gold and one of copper. In the United States coin, there is also one-twelfth of alloy, but this may be composed of any proportions of copper and silver, that may be convenient, provided the latter be not in excess.

*Standard silver*, in Great Britain, contains one-twelfth of copper; in the United States, the proportion is one-ninth. The various silver utensils and ornaments, are generally made of this kind of silver.

*Potassium* and *sodium* form alloys, differing in their properties, according to the proportions in which they are combined. A small quantity of sodium renders potassium fluid at ordinary temperatures. Eight parts of potassium and one of sodium, form a compound that swims on naphtha, and is fluid at  $60^{\circ}$ .

Mercury amalgamates rapidly with a number of metals. Gold and silver are rendered very brittle, when combined with a very small quantity of this metal.

Various compositions have been used, to form an amalgam for the electrical machine; one part of zinc with an equal weight of tin and two parts of mercury, answer very well for this purpose.

When mercury is poured on tin foil, it spreads rapidly over its surface, forming an amalgam which is used in making looking-glasses.

An amalgam composed of one part of tin, one of lead, two of bismuth, and four of mercury, is employed for silvering the inside of hollow globes of glass.

If one part of mercury be added to the *fusible metal*, an amalgam will be obtained, which becomes soft at  $162^{\circ}$ , and

quite fluid at  $170^{\circ}$ . This is the best compound for making *fusi-  
ble spoons*.

Iron is coated with tin, by dipping it in a melted alloy, consisting of ten parts of tin and one of copper. The copper prevents too large a quantity of tin from adhering to the iron, and the melted metal should be covered with a little tallow to prevent it from oxydating. It is in this manner, that *tinned iron* is prepared, consisting of thin sheet iron, coated with tin.

Copper vessels are coated with tin, by rubbing them over with muriate of ammonia, first throwing a little powdered rosin on them, to prevent the copper from acquiring any crust of oxyde, and spreading a little melted tin on their surface, after exposing them to heat.

Copper may be coated with silver, by rubbing it with the following mixture made into a paste with water, after boiling it in water with cream of tartar and alum; one part of the bi-chloride of mercury, four of silver, precipitated from a solution of the nitrate of copper, sixteen of common salt, and an equal quantity of the muriate of ammonia. The copper must then be exposed to a red heat to drive off the mercury, after which the coating of metallic silver may be polished.

Silver, copper, and brass, are easily gilded, by an amalgam composed of eight parts of mercury and one of gold, prepared by heating a mixture of the above metals in these proportions, till the gold is completely dissolved, and pouring it immediately into cold water, to prevent any loss of mercury. The metal to be gilded, is previously washed with a dilute solution of the nitrate of mercury, containing an excess of acid, and the amalgam laid over, as uniformly as possible, with a brush made of brass wire. By exposing it to heat in a furnace, or over an open fire, the mercury is dissipated, and the coating of gold left in combination with the metal, after which it may be polished.

Steel may be gilded, by dipping it into the ethereal solution of gold, prepared in the manner described in the chapter on gold.

The gold powder, used in painting, is prepared by exposing an amalgam of gold and mercury to heat, till all the mercury is volatilised, or by triturating gold leaf with a solution of gum, and washing away the gum afterwards with a large quantity of water.

## PART III.

### OF ANIMAL SUBSTANCES.\*

---

#### SECTION I.—FIBRINE.

FIBRINE may be easily procured, in a pure form, by pouring off the serum from coagulated blood, and washing the crassamentum in a linen or cotton bag, with a large quantity of water, till all the coloring matter is carried away; the fibrine remains of a white color, with a tint of yellow, which is more apparent in some kinds of fibrine than others. It may be obtained in a state of sufficient purity for all ordinary purposes, by cutting muscular fibre to small pieces, and washing it repeatedly with water.

Put some fibrine into alcohol of the specific gravity of 0.80, and set it aside for several weeks; the fibrine is slowly converted into a fatty matter, which is dissolved by the alcohol, and may be precipitated again from its solution by water.

Put some muscular fibre into a wooden box with holes in it, and place it where it will be exposed to a stream of water for two or three weeks; a change will take place similar to what is effected by the alcohol, the fibrine being converted into adipocere.

Fibrine may be easily converted into a fatty matter, by the action of diluted nitric acid. Mix three or four ounces of nitric acid with one and a half times their weight of water, and expose the liquid to heat over a chauffer, with two or three parts of muscular fibre. Nitrogen gas will be disengaged, not mixed with any protoxyde of nitrogen at first, and appearing to arise

solely from the decomposition of the fibrine. In a short time, however, the acid itself will begin to be decomposed, a large quantity of fatty matter will be formed, and protoxyde of nitrogen and vapors of nitrous acid will be evolved. The nitrogen disengaged in the first stage of the process, may be collected, if required, by heating the mixture

\* Having already exceeded the limits originally prescribed for this volume, I am compelled to be very brief, in the examination of the subjects that are yet to be noticed.



in a glass retort, the beak of which is introduced under the surface of water in a pneumatic trough, arranged with jars in the usual manner.

Pour some concentrated acetic acid over fibrine, in a Florence flask, and allow them to remain mixed together for some time; the fibrine gradually softens, and on exposing it to a gentle heat, it will be dissolved, and a gelatinous mass obtained, which is soluble in water; the fibrine appears, however, to be partially decomposed, as a little nitrogen gas is disengaged during the solution of the jelly in water. Exp.

Put some fibrine into a solution of potash or soda, at ordinary temperatures, and allow them to remain mixed together. The fibrine will be slowly dissolved, and be precipitated again from the solution, by neutralising the alkali with an acid. Exp.

## SECTION II.—ALBUMEN.

Put a piece of blue test-paper into the white of an egg, and it will soon turn green; the albumen and water, of which it is almost entirely composed, containing a portion of free soda. A similar experiment may be made with the serum of the blood, which also is composed principally of water and albumen, with a little free soda. Exp.

Expose some of the white of an egg, or the serum of the blood, to a temperature not exceeding  $160^{\circ}$ ; in a short time it will be converted into a firm coagulum. Exp.

Mix some albumen with a considerable quantity of cold water, and expose it afterwards to heat; the albumen will be coagulated and the water will become opaque, even though the liquid should not contain more than one-thousandth part of albumen. Exp.

From the property which albumen has of being miscible with cold water, while it is coagulated by hot water, it is used in many chemical operations for clarifying saline solutions which do not coagulate it at natural temperatures. For this purpose, the albumen is added to the liquid to be clarified, while cold, and on exposing it to heat, the albumen coagulates slowly as the temperature increases, forming a kind of net-work which envelopes the dust and any other extraneous matters which the liquid may contain, and collecting a scum at the surface, which is easily removed.

Put the white of an egg into a Florence flask, and pour an ounce or two of sulphuric acid upon it; the albumen will be immediately coagulated, and on applying a gentle heat, it will be completely dissolved; it is partly decomposed, Exp.

however, at the same time, and the liquid assumes a very dark color, having a purple tint. When the heat was very carefully applied, Dr Hope found that the liquid became of a very beautiful red color, which may be regarded as one of the characteristic properties of albumen.

Mix the white of an egg with a considerable quantity of water, and pour the liquid into a number of glasses; then add  
 Exp. a solution of the bi-chloride of mercury to one, a solution of the subacetate of lead to another, and a solution of several other metallic salts to the rest. The albumen will decompose the solutions of a great number of metallic salts, combining with the metallic oxide and forming an insoluble compound which is precipitated. The precipitate that is thrown down from the bi-chloride of mercury, is regarded as a compound of the protochloride of mercury and albumen; smaller quantities of albumen may be detected in this manner, than in any other way; a single drop of a saturated solution of the bi-chloride rendering water turbid, when it contains only a two-thousandth part of albumen in solution.

Add an infusion of galls to a similar solution of albumen in water, a copious precipitate will be immediately thrown down, composed of albumen and tannin.

Expose an ounce or two of the serum of the blood to a very gentle heat, till it is coagulated; cut the coagulum into small pieces and put it on a filter. A watery fluid (usually termed the serosity) will exude, which may be separated easily, by pouring more water on the coagulum, and which contains in solution the greater part of the salts of the blood; muriatic acid may be easily detected in it, by a solution of the nitrate of silver, which gives a copious, white, curdy precipitate with this liquid.

#### SECTION.—III. GELATINE.

Dissolve one part of dry gelatine in 100 parts of hot water, and a liquid will be obtained, which will become a tremulous jelly when cold. Add an infusion of galls to a solution

Exp. of gelatine in water; a copious precipitate will be immediately thrown down, composed of tannin and gelatine, similar in its composition to leather, and on this principle, the processes of the tanner are based.

Add a solution of the bi-chloride of mercury to a solution of gelatine and albumen, as long as any precipitation takes place; the albumen will be precipitated and the gelatine will  
 Exp. remain in solution; it may be removed afterwards by a solution of tannin. If the tannin be added to the liquid containing the albumen and gelatine, before precipitating the former

by bi-chloride of mercury, both the albumen and the gelatine will be thrown down in combination with it.

For all ordinary experiments where gelatine is required, common glue will be found to do very well; where it is necessary to have a purer gelatine, isinglass should be employed. It may be procured, by boiling animal substances in water, and evaporating to dryness.

#### SECTION IV.—OSMAZOME.\*

This proximate animal principle may be obtained, by macerating muscular fibre in cold water, evaporating the liquid procured in this manner to the consistence of an extract, after separating the coagulated albumen, and digesting it in alcohol. The water dissolves the albumen, osmazome, and saline matter which are associated along with fibrine in the muscular fibre; the albumen is coagulated by the heat, and the osmazome is separated by the alcohol from the extract obtained on evaporation, being soluble both in water and alcohol. Its solution does not gelatinize on cooling. It has the smell and taste of the juice of meat.

#### SECTION V.—MISCELLANEOUS EXPERIMENTS, ILLUSTRATING THE METHOD OF EXAMINING THE MOST IMPORTANT VARIETIES OF CALCULI.

Cut one or two *biliary* calculi into small pieces, triturate them afterwards in a mortar, and boil them for ten minutes in a test-tube or small flask, with twenty times their bulk of alcohol. On filtering the solution, a clear and colorless liquid is usually obtained, which deposits pearly looking crystals of *cholesterine*, as it cools. If the solution be very strong, a layer of *cholesterine* will be left upon the paper by the alcohol as it passes through, and on drying the filter and bending it backwards, it may be detached in the solid form, still retaining the figure of the filter on which it had collected, and having the appearance of a fine membrane, with a pearly lustre.

In preparing *cholesterine* from *biliary* calculi, it is necessary to filter the solution as speedily as possible, as this substance is

\* This term appears to be derived from two Greek words, signifying, literally, the *odor* or *smell of bodies*.

very sparingly soluble in cold alcohol, and a great portion of it is deposited on the filter in operating with small quantities, if this precaution is not attended to.

Reduce a portion of a *uric acid calculus*\* to powder, and digest it in a diluted solution of potash. Urate of potash will be formed and remain in solution, while any phosphate of lime, or phosphate of ammonia and magnesia, that may have been mixed with it, will remain undissolved.

Filter the solution of urate of potash, prepared in the manner described in the preceding paragraph, and add acetic acid to it as long as any precipitation takes place. The acetic acid combines with the potash and uric acid is precipitated in very small crystals.

Mix some uric and nitric acids, and expose the mixture to a gentle heat till it becomes dry, taking care not to allow the temperature to become so high as to decompose the products; a solid mass is obtained of a deep red color. When a very small quantity of materials is employed, so as to leave only a very thin pellicle on the surface of the glass on which it is heated, it has a vivid pink color.

Expose a small piece of a uric acid calculus to heat in a platina spoon, held in the flame of a spirit lamp; it soon blackens and emits a fetid odor, and in a short time it entirely disappears, when composed of nothing but uric acid; in general, however, an ash colored matter remains, composed of some of the saline substances, that form a number of other calculi, and occur in variable proportions in the uric acid calculus.

The *urate of ammonia calculus* is, comparatively, rare; and till lately, the formation of calculi of this kind was disputed. It may be easily distinguished from the uric acid calculus, however, when not mixed with a large quantity of uric acid, by its greater solubility in water, its clay color, the ammoniacal odor which is exhaled when gently heated with a strong solution of potash, and its solubility in solutions of the carbonates of potash and soda, which do not dissolve pure uric acid.

The *oxalate of lime* or *mulberry calculus* is, in general, easily known by its rough tuberculated surface, and its chemical composition may be easily proved by a few simple experiments.

Expose a portion of this calculus to heat over a spirit lamp, holding it in the flame till nothing remains but a gray colored ash; it will become black at first, and the residuum will be found to be composed of carbonate of lime, dissolving

\* When this acid has not accumulated in mass, but is merely suspended in the urine, it is called *red sand*.



with effervescence in muriatic acid, and forming a solution of muriate of lime.

Place another portion of this calculus on a piece of charcoal, and expose it to a strong heat before the flame of the blowpipe; nothing will remain but caustic lime, which will slake in the usual manner, and render turmeric paper brown, Exp. and paper tinged with the blue infusion of cabbage, green, on rubbing it on them with a little water.

If it be required to show the presence of the oxalic acid in the calculus, it must be reduced to a fine powder, and digested with a very gentle heat in diluted sulphuric acid; sulphate of lime will be formed, and oxalic acid will be found in solution. In general, however, this process may be dispensed with.

Reduce a portion of the *phosphate of lime* or *bone earth calculus* to powder, and digest it in diluted nitric or muriatic acid. In a short time it will be completely dissolved; Exp. and if the acid be neutralised by an alkali, it will again be precipitated.

Expose another portion of this calculus to heat, before the flame of the blowpipe; it will soon become black from the decomposition of a little animal matter with which it is always mixed, and the separation of part of the carbon; Exp. this will soon be burnt off, however, and nothing will remain but the phosphate of lime, of a pure white color, which does not suffer any farther change, unless it be exposed to a very intense heat, by which it may be fused.

Digest another portion of this calculus in a solution of potash, after reducing it to powder; if it consist solely of phosphate of lime, no reaction will take place, phosphate of lime being insoluble in a solution of potash; but if any uric acid Exp. should have been mixed with it, urate of potash will be found in the solution, from which the uric acid may be precipitated by another acid.

Uric acid may be detected, also, in calculi composed principally of phosphate of lime, by evaporating its solution in nitric acid to dryness, when the characteristic red tint will appear, though only a very minute quantity of this substance should have been present in the calculus. There are few urinary calculi, which do not give traces of uric acid when treated in this manner, whatever may be the nature of the ingredient of which they are principally composed.

When the calculus is supposed to consist of *phosphate of ammonia* and *magnesia*, this may be easily ascertained, by heating a portion in a solution of caustic potash, after reducing it to powder; the ammonia will be disengaged, and the phosphoric



acid will unite with the potash; and by treating another portion with cold acetic acid, the phosphate of ammonia and magnesia will be dissolved, leaving the greater portion of the phosphate of lime that may have been mixed with it. This variety is sometimes denominated *white sand*.

When the calculus is composed of a mixture of phosphate of ammonia and magnesia and phosphate of lime, it fuses readily before the blowpipe, and hence, it is usually termed the *fusible calculus*.

The *cystic oxyde calculus* is easily distinguished from all the preceding varieties of calculi, by its solubility in solutions of acids, alkalis, alkaline carbonates, and in lime water; it is also completely decomposed by heat, being composed of the usual elements of animal matter.

The *xanthic oxyde calculus* is so rare, that few students can expect to have an opportunity of operating with it. It is distinguished by the lemon-yellow colored residuum that is left, on exposing its solution in nitric acid to heat. The *fibrinous calculus* is still more rare, and is composed entirely of fibrine.

From the above experiments, although briefly stated, the student may perceive, that not a little light has been thrown by chemistry on the treatment of calculous affections. He can also discern the fitness of alkaline remedies for the prevention and cure of *uric acid* or *red sand calculus*, and the propriety of acids in the treatment of what is called the *white sand calculus*.

## PART IV.

### SECTION I.—ANIMAL AND VEGETABLE ANALYSIS. THE MODERN PROCESS ILLUSTRATED BY EXPERIMENTS.

THOUGH the exact analysis of animal and vegetable substances is a process far too complicated for the beginner, and requires a knowledge and dexterity of manipulation that can be acquired only by long practice, it may be interesting for him to imitate the process, that he may see a practical illustration of the principle on which it depends. The two following experiments will be sufficient for this purpose, and they may be performed easily by the beginner.

Mix five grains of sugar with fifty grains of the peroxyde of copper, prepared by decomposing the nitrate of copper by heat. Put this mixture into a glass tube about six or eight inches long, and a third of an inch in diameter, and join to the extremity a smaller tube, bent at the outer end. Join these together, accurately, with plaster of Paris and a little mucilage, securing the whole with some stout thread. The extremity of the tube containing the mixture is then to be coated with clay, and wrapped round with harpsichord wire, after which it may be put into a chauffer with an aperture at the side, placing the extremity of the small tube under mercury, and holding over it another tube closed at one end, previously filled with mercury, that the disengaged gas may be collected in the same manner as in a jar over the mercurial trough. This tube need not be above ten or twelve inches long, nor more than half an inch in diameter, so that the operator may hold it easily in one hand, while he directs the extremity of the tube from which the gas is escaping with the other. He should also have a tin plate to protect his hand from the heat of the chauffer, and a chimney to increase the temperature, if the chauffer should not burn well.

Exp.

If it be required to collect more than one tube of gas, the other tubes intended to be used, should be filled with mercury before commencing the process, and arranged in another cup containing this fluid, supporting them at the top, within the ring of a retort stand. The operator will thus be enabled to

remove each tube, as it is filled with gas, into this cup, closing it completely with his thumb as he transfers it from the one to the other, and replacing it by a tube full of mercury, which must be taken in the same manner to the first cup, to be filled with gas, as in the preceding instance.

The pressure to be overcome by the disengaged gas, in rising through the mercury, being about thirteen times greater than it meets with in passing through a column of water of the same height, the two tubes which are joined together must be luted with great care. The student, who has attended to the method of adjusting tube apparatus, may save himself the trouble of attaching another tube to the one containing the mixture, by bending and drawing out the extremity till it is obtained of a form similar to that of the two tubes, when joined and luted; this must not be done, however, till the mixture has been introduced, as it would be troublesome to pass it through the small aperture at the extremity, which need not exceed the sixteenth part of an inch in diameter.

When the gas ceases to come, a few drops of a solution of caustic potash may be introduced into one of the tubes filled with gas, under the surface of the mercury, by means of a pipette; in a short time it will be completely absorbed, if all the atmospheric air shall have been allowed to escape before collecting any gas, nothing being produced by the decomposition, but carbonic acid and water, and the latter being deposited in the tube; the quantity of water is ascertained, in delicate experiments, by placing fragments of chloride of calcium within that part of the apparatus through which the carbonic acid is made to pass, this substance having a great affinity for water, and retaining it in combination; the weight which it gains during the operation, denotes the quantity of water produced.

The theory of the process will be readily understood, from what has been stated in a former part of this work. The sugar, being composed of carbon, hydrogen, and oxygen, the only products are carbonic acid and water, the peroxyde of copper supplying the additional quantity of oxygen, necessary to convert all the carbon and hydrogen into carbonic acid and water. By weighing, then, the carbonic acid, we may ascertain the quantity of carbon, every twenty-two parts indicating the presence of six of this element; nine parts of water denote the presence of one of hydrogen; and by comparing the quantity of oxygen which these products contain, with the weight of the sugar employed, and the quantity of carbon and hydrogen which it has been found to contain, we may easily ascertain the proportion of oxygen; the latter may be found out also by comparing the

quantity of oxygen in the carbonic acid and water with the loss of weight which the peroxyde has sustained.

For the second experiment alluded to, the beginner may mix, intimately, five grains of dried gelatine with fifty grains of peroxyde of copper, covering the mixture with an additional quantity of the peroxyde after putting it into the tube, and then proceeding in the manner directed in the preceding experiment. It will be better, however, to heat that part of the tube containing the peroxyde of copper put over the mixture, in the flame of a spirit lamp, before putting the coated extremity into the chauffer, the peroxyde, heated in this manner, decomposing any oily matter or impure carbonate of ammonia that is sometimes disengaged, before the rest of the peroxyde is heated sufficiently to afford it oxygen with facility. In this process, carbonic acid, water, and nitrogen are the only products, the two first being formed in the same manner as in the preceding process, while the nitrogen is disengaged and collected in the tube, along with the carbonic acid. On introducing a solution of caustic potash, the whole of the carbonic acid gas will be absorbed, and the residual gas will be found to be nitrogen, being unflammable, not supporting combustion, and giving no precipitate with lime water.

When several experiments of this kind are to be performed, a copper tube will be found more convenient, which may be conveniently heated by a number of spirit lamps; but for the above experiments, a coated glass tube will be found quite sufficient. Green glass tubes should be preferred, as they are not so apt to soften and give way when exposed to a red heat.

## SECTION II.—OF MINERAL WATERS AND THEIR ANALYSIS.

Waters are usually denominated *mineral*, when they have acquired a peculiar flavor or taste, owing to the soil or strata through which they have passed; and are thence regarded as unfit for domestic uses. They, of course, differ from all the varieties of what are termed *pure* waters, such as *rain*, *snow*, *spring*, *well*, or *river* water. The purest water, obtained apart from distillation, is that which is collected, directly, in clean vessels, as it is poured out from the clouds. If it be received from a roof, it is less pure, and the degree of purity will vary according to existing circumstances.

Freshly fallen and quickly melted snow, also, affords a water of considerable purity; but this, as well as rain water, contains carbonic acid, and is said to be very rich in oxygen.



When water falls to the ground and percolates through the soil, its solvent power enables it to take up a variety of substances; but when these are small in quantity and not calculated to impair the taste and other sensible properties, the water is still affirmed to be comparatively pure. Of this nature, is the ordinary *spring, well, and river* water.

Spring water is sometimes very pure. This is the case, especially, when the fluid has found its way, by a sort of filtration, through quartz, granite, and the like; it is then almost entirely free of taste or odor. But when it percolates through the crevices of limestone or gypsum, it always acquires that peculiar property, to which the term *hardness* has been applied, owing to the fact, of its power to dissolve a small portion of those substances. This property of hardness is well known in domestic economy, and the use of potash, to remedy the difficulty, is also familiar to many persons. We test this property by adding the tincture of soap, which instantly produces a curdy precipitate. This arises from the union of the lime with the oleic and margaric acids of the soap, whence results an insoluble compound; while the potash of the soap joins the acid, previously united to the lime, and the result is a soluble compound. When carbonate of lime causes the hardness, (or rather the bi-carbonate,) we may throw down the lime, by boiling. The carbonic acid is thus expelled, so as to render the lime insoluble and it is necessarily precipitated. But if the hardness be caused by sulphate of lime, then the carbonate of potash or soda will be a proper addition, and the water will be made soft by the precipitation of the lime, in the form of oleate and margarate of lime.

Sometimes we resort to distillation, to free even the purest natural water from substances that ordinarily exist in it; and this is the most certain method we can adopt.

Various divisions have been made by writers, in treating of mineral waters, but we regard the following as the best, viz. the *acidulous*, the *alkaline*, the *chalybeate*, the *sulphurous*, the *saline*, and the *silicious waters*.

1. *Of Acidulous Mineral Waters.* These are found in nearly all countries, and are dependent, for their peculiar properties, on the presence of free carbonic acid, and hence, the facility with which they have been imitated by art. Every one has heard of the sparkling waters of the Seltzer springs, and no one is ignorant of the agreeable beverage sold in our apothecary shops, under the title of artificial mineral or Seltzer water. In various sections of the United States, waters are found, that have a similar impregnation.

All carbonated waters give a red tint to litmus paper, when



taken fresh from the spring or other source; but this effect will not be seen, after the water has been boiled for a short time, owing to the expulsion of the acid by the repelling power of heat. So, also, if we add lime water to these liquids, a turbidness is apparent, resulting from the union of the carbonic acid of the mineral water with the lime, in the form of an insoluble carbonate. Exp.

In addition to carbonic acid, these waters often contain lime, magnesia, and iron, which are held in solution, by an excess of carbonic acid, giving rise to the form which is called bi-carbonate and which is always soluble. If a portion of these waters be boiled, so as to expel a part of the carbonic acid, there will be a precipitate formed, of carbonate of iron, which is insoluble. Exp.

We can determine how much carbonic acid a mineral water contains, if the acid be in a free state, by putting a given quantity into a common flask, furnished with a bent tube, the outer end of which is passed under a graduated jar, filled with and inverted over mercury. If heat be applied to the flask, it will expel all the carbonic acid, and it can easily be seen, what quantity has entered the jar.

By these simple rules, any one can determine, with facility, not only the fact that free carbonic acid is present in a mineral water, but also, with considerable accuracy, the actual quantity.

2. *Of Alkaline Waters.* If we dissolve soda or its carbonate, in water, the solution may be denominated an alkaline water. Such is, in fact, the nature of the real soda water, kept in some of our apothecary shops; though it often occurs, that a great deal of artificial mineral water is sold as soda water, in which there is not a particle of alkaline matter.

The springs which are known to furnish alkaline waters, are comparatively rare. They contain carbonate of soda, and carbonic acid in a free state, with scarcely any admixture of earthy substances. It is affirmed, also, that these waters usually contain protoxyde of iron, sulphuretted hydrogen gas, and muriate of soda.

All waters that merit the title of alkaline, will give the usual alkaline reaction. Thus they will change to brown, the infusion or tincture of turmeric, and to green, the blue infusions. Add, therefore, to a diluted infusion of blue cabbage, some of the water under examination; if it be an alkaline water, the color of the infusion will be changed to green. Exp.

3. *Of Chalybeate Waters.* These, as the name imports, are all those waters which are characterised by the presence of iron. They have a strong, styptic taste, and strike a black

color with the infusion of nutgalls. Occasionally, the iron is combined with muriatic or sulphuric acid, but more frequently it exists in the form of carbonate, or rather bi-carbonate of the protoxyde. If the water be taken from its source and left exposed to the open air, the protoxyde of iron acquires additional oxygen, and is changed to the peroxyde, which subsides as an ochreous deposit. This substance, which is of a reddish-yellow color, is seen in the immediate vicinity of all chalybeate springs. A portion of the water is carried, by an overflow, on the surrounding soil, and there the iron is changed from a soluble protoxyde, to an insoluble peroxyde.

It is always desirable to determine the quantity of iron contained in a mineral water; and this may be done with sufficient precision, by evaporating a given portion of the water, almost to dryness, and then converting the iron into peroxyde, by means of nitric acid. The nitrate of the peroxyde, thus formed, is then to be decomposed by the addition of an alkali, which uniting with the acid, causes the peroxyde to be detached, and to fall down as an insoluble matter, which may be weighed, after having been dried.

Should lime and magnesia also be present in the water, they may be separated by the following processes. Add to a portion of the water, the oxalate of ammonia, and the oxalic acid  
 Exp. will seize the lime, forming an insoluble oxalate of lime. Let this be collected and exposed to a red heat; the oxalic acid will be driven off and pure lime remain, which can be easily weighed.

To the same water, after filtration, add an excess of carbonate of ammonia, in solution, and then the phosphate of soda, when the magnesia will be precipitated in the form of  
 Exp. the ammoniaco-phosphate of magnesia. Heat this precipitate to redness, and every 100 parts of it are equivalent to 37 of pure magnesia.

No mineral waters are more common than those denominated chalybeate; but when a patient is too infirm to visit a spring for the purpose of drinking the waters, a very good substitute may be formed, by the use of the following formula, furnished by Dr Ure. Take three grains of sulphate of iron, sixty one of bi-carbonate of potash and a quart of cool water, and put them into a vessel, having a tight stopper. Let them be freely shaken, and in a few hours, the water will be impregnated with bi-carbonate of iron, as the result of double decomposition. I will not say, that a water thus prepared, although more strictly medicinal than the native fluid, will have equally salutary effects, for it cannot be doubted, that much of the happy results of a visit to

a spring, depends on the agency of new associations, such as the company, the scenery, and a variety of other items, which every one knows to be peculiar to such locations.

4. *Of Sulphurous Waters.* The sensible presence of sulphuretted hydrogen gas, is always sufficient to determine that a mineral water belongs to this class. It cannot be mistaken by any one, whose olfactories are not in a state of disease or strange perversion. These waters are very common, and although they are often beneficial, are, at the same time, extremely offensive. We have said, already, that a few grains of the carbonate of lead, (the white lead of painters,) added to a water containing the least quantity of sulphuretted hydrogen gas, will indicate its presence, most certainly. A brown or blackish precipitate, of sulphuret of lead, will be instantly formed. We expel this gas with great ease, by the process of boiling, and if the whole of the gas, as it is evolved, be received into a solution of the acetate of lead, we can determine its proportion, by weighing the sulphuret of lead, thus formed. Every 120 parts of the sulphuret, denote the presence of 17, or one equivalent of sulphuretted hydrogen gas.

5. *Of Saline Waters.* All waters which do not strictly belong to either of the above classes, are embraced in the denomination of *saline*. These contain the sulphates, muriates, carbonates of lime, magnesia, and soda. Sometimes potash is present, and, rarely, small portions of lithia. Hydriodic and hydrobromic acids have also been found in these waters, not only in Europe, but in this country, as in the Kenhawa springs, the Saratoga waters, &c. These acids are usually combined with potash or magnesia.

It is always desirable, to ascertain the nature of the acid, in the first instance, that enters into the formation of the saline matter. If muriatic acid be present, nitrate of silver will indicate it, by giving a white precipitate of chloride or muriate of silver. Exp. The muriate of barytes will detect sulphuric acid, giving an insoluble precipitate of sulphate of barytes; and if a carbonate be present, the precipitate occasioned by either of the above tests, will contain carbonate of silver or barytes.

The presence of lime or magnesia may be ascertained, by the modes already pointed out. If potash be held in solution, it will be made manifest by the yellow precipitate, resulting from the addition of a few drops of the muriate or chloride of platina. Exp. To detect soda, evaporate a portion of the water to dryness, remove the deliquescent salts by means of alcohol, and all that is insoluble in that fluid should be taken up by means of water. This latter solution contains the soda,

which is obtained in crystals, by spontaneous evaporation, and gives to flame a rich yellow color.

To detect hydriodic acid, we are obliged to resort to the well known test of iodine. The water under examination is first evaporated to dryness, and the soluble matter of the residuum is dissolved in two or three drachms of a cold solution of starch,

*Exp.* after which a little strong sulphuric acid must be gradually added. The latter detaches the potash from its connexion with the hydriodic acid, which is then in a proper state to be acted on by the starch. The iodine of the hydriodic acid combines with the starch, forming ioduret of starch, which has a deep blue color.

Hydrobromic acid is detected, by virtue of the action of chlorine. When chlorine gas is added to a water containing the most minute quantity of bromine, in the shape of hydrobromic acid, (or rather as a hydrobromate of potash,) it detaches the

*Exp.* hydrogen from this acid, and sets the bromine free. This liberation is evinced, by the orange-yellow color imparted to the liquid; and if the solution be heated to the boiling point, the red vapors of bromine will be expelled, and may be condensed, by being conducted into a tube surrounded with ice.

Having in this manner ascertained the nature of the saline ingredients, it is important to determine, with tolerable accuracy, their total, as well as their respective quantities. M. Kirwan has furnished a rule, by which to ascertain the proportion of saline ingredients in a water of known specific gravity, and it is as follows; subtract the specific gravity of pure water, from that of the water examined, and multiply the remainder by 1.4. The product is equal to the saline contents, in a quantity of the water denoted by the number employed to indicate the specific gravity of distilled water. Thus, suppose the specific gravity of the water to be examined = 1.079, and that of pure water = 1.000; then  $79 \times 1.4 = 110.6 =$  saline contents in 1000 of the mineral water.

We may also determine the whole amount of saline matters, by evaporating a pint of water to dryness, heating to low redness and weighing the residue.

To make an exact analysis, a given quantity of the water should be concentrated in an evaporating dish, as far as practicable, without causing precipitation, or crystallisation, and the residual liquor should then be divided into two parts. From one portion, throw down the sulphuric and carbonic acids, by the addition of nitrate of barytes; then filter, and to the clear



liquor, add nitrate of silver, which will detach the muriatic acid, in form of muriate or chloride of silver. The mixed sulphate and carbonate is then exposed to a low red heat <sup>Exp.</sup> and weighed, and the latter is dissolved by diluted muriatic acid, and its quantity determined by weighing the sulphate that is left. The chloride of silver, of which 146 parts correspond to 37 of muriatic acid, is fused in a platina spoon or crucible, to render it free of moisture. The other half of the concentrated mineral water is next to be examined, to determine the quantity of lime and magnesia, in the way already pointed out. Having thus determined the weight of each of the fixed ingredients, excepting the soda, the apparent loss or difference in weight, necessarily gives the quantity of that alkali.

Such are the leading features in the process for examining saline waters, and the reader is referred to Henry, Ure, and others, for more minute information.

6. *Of Silicious Waters.* These are exceedingly rare, and have never, so far as I am informed, been found in this country. Although silica is the predominating ingredient, it is always united with soda and alumine, in various states of combination.

### SECTION III.—OF COLORING MATTERS, AND OF THE PREPARATION OF TESTS AND TEST-PAPERS.

To prepare the colored test-papers which have been so frequently referred to in this work, a red cabbage must be cut into slices and boiled for half an hour in a pint or two of water, adding more water, from time to time, if it should be required, to prevent any of the vegetable matter from being decomposed by the heat. The liquid is then to be decanted from the fibrous matter that remains, and concentrated by evaporation, till it is reduced to five or six ounces by measure, pouring it afterwards into a flat plate or bason, and steeping in it pieces of unsized writing paper, (common printing paper does very well,) taking care to moisten every piece thoroughly, by drawing it backwards and forwards several times through the liquid, before they are laid above one another. In half an hour they may be taken out to dry, hanging them on a piece of cord, drawn tight between two nails on the opposite sides of a room, where they will not be exposed to dust or acid fumes. The paper may be cut into pieces about the size of an octavo page, before they are immersed in the solution, and each of them may be divided across into thirty or forty pieces, putting them into a tin case or any other convenient vessel, that they may be always at hand, as they are



continually required in a great variety of operations. A sufficient quantity may be made at once in this manner, to serve for a year or two, even where a great number of experiments is performed.

Paper dyed blue with the coloring matter of the cabbage, is used for indicating the presence of acids and alkalis, being turned red by the former, and green by the latter. It is sufficiently delicate for all ordinary experiments, and is preferable to litmus paper, which is not affected by alkalis, unless previously reddened by an acid, when the original blue is restored.

When the coloring matter of the cabbage is required in solution, it may be easily obtained by infusion in water. It begins to pass into a state of putrefaction in a few days, however, exhaling a very disagreeable smell, and the color fading at the same time. The concentrated solution may be kept for a long time when put in bottles which have been completely filled with it, after mixing every ounce or two with a drachm of sulphuric acid. Mr Faraday states, that it may be preserved in this manner for a year.\*

*Litmus* and *turmeric test-papers* may be prepared, by a similar process. The litmus must be reduced to fine powder and boiled with water, to procure a solution fit for the purpose; turmeric water may be obtained sufficiently strong, by pouring boiling water over this substance in powder. Litmus paper is used as a test for the presence of acids, which immediately redden it, and as a test for alkalis, when it has been previously reddened by an acid. Turmeric paper is changed to a reddish-brown color, by alkalis and alkaline earths, but is not affected by the acids; it may be used, however, as a test of acids when it has been turned brown by an alkali, the acid combining with the alkali and restoring its original yellow color. It cannot be relied on, however, as a test for the presence of alkalis, a number of compound salts and other substances affecting it in the same manner as these bodies, as Mr Faraday has pointed out.

### *Experiments with Coloring Matter.*

Put a small quantity of a strong solution of the muriate of tin into seven or eight ounces of a solution of litmus, and add a solution of potash as long as it produces any precipitate.

Exp. The potash combines with the muriatic acid, forming muriate of potash which remains in solution, and the oxyde of

\* The blue paper in which loaf sugar is usually enclosed, answers very well as a test for acids, and can be had at all times.

tin, uniting with the litmus, forms an insoluble compound, which is immediately precipitated. It is in this manner, that the pigments called *lakes*, are usually prepared.

Make a similar experiment with solutions of alum (sulphate of alumine and potash) and litmus; on adding the solution of potash, the alkali combines with the sulphuric acid, previously in combination with the alumine, which is thrown down in combination with the litmus. Exp.

As the oxyde of tin and alumine can combine only with a certain quantity of coloring matter, if the whole of the litmus be not precipitated at first, and the liquid left transparent and colorless, by adding an additional quantity of muriate of tin or alum to the solution, and precipitating as before by potash, all of it may be thrown down.

Make a strong solution of cochineal in water, and precipitate the coloring matter in combination with alumine or oxyde of tin, by a solution of potash. A rich colored lake is immediately precipitated, well known by the name of *carmine*. Exp.

Acetate of alumine is generally preferred to alum, in processes where alumine is to be separated in combination with coloring matter, as the acetic acid that is disengaged, is not so liable to react upon cloth or coloring matter, as sulphuric acid, which often proves troublesome when alum is used. It may be easily obtained, by the process mentioned, when treating of the salts of alumine.

Digest some indigo, reduced to powder for an hour or two, in sulphuric acid; a solution will be obtained which has a greenish-yellow color at first, and passes to a deep blue when it has been kept for several hours. Dilute the solution with a large quantity of water, and put a piece of silk, linen, or cotton into it, taking it out immediately and allowing it to dry. Part of the indigo will combine with the cloth, and it will be found to be dyed blue, the shade being more or less deep, according to the strength of the solution. Exp.

Dissolve some sulphuret of arsenic in a hot solution of potash, and digest indigo reduced to a fine powder in the liquid. The metallic sulphuret attracts oxygen from the indigo, when it will become of a green color, and be immediately dissolved. Put a piece of cloth into the solution, and take it out to dry, exposing it freely to the air. The cloth will combine with a portion of deoxydated indigo immediately, and appear of a green color, when taken out of the liquid; it soon attracts oxygen from the air, however, the indigo recovering its original color, and in this manner cloth is usually dyed blue. Exp.

Expose some indigo to heat over a chauffer in a glass vessel; it is sublimed at the temperature of  $550^{\circ}$ , a reddish-violet colored vapor being formed, which bears a considerable resemblance in its appearance to iodine, and condenses in minute crystalline grains.

Boil half an ounce of powdered Brazil wood, for a quarter of an hour, in eight or ten ounces of water; filter the solution of the coloring matter obtained in this manner through a double linen cloth, and dip a piece of cotton cloth into the liquid. It immediately combines with part of the coloring matter, but the color is neither bright nor permanent.

Take another piece of the same cloth, pass it through a dilute solution of the muriate of tin, and allow it to dry. Then put it into the same solution of the coloring matter of the Brazil wood;

the compound of the cloth and the oxyde of tin that is formed in the preceding part of the process, attracts coloring matter from the solution and is immediately dyed red; the color is also much brighter and more permanent than before, the oxyde of tin retaining it more strongly in combination with the cloth, by the great attraction which it has for both; when a common metallic oxyde or earth is used for this purpose, it is in general termed a *mordant* or *basis*.

Similar experiments may be made with cochineal, madder, logwood, and safflower, all of which are red dyes, and termed *adjective colors*, as they require the intervention of some basis to

render them permanent, having a feeble affinity for different kinds of cloth. When the affinity of coloring matter for cloth is so great that no basis is required to effect a permanent combination, it is usually termed a *substantive color*, of which indigo is a very good example.

Cloth may be dyed yellow, by proceeding in the same manner with solutions of the coloring matter of saffron, fustic, quercitron bark, or hickory. There are several other substances used for dyeing yellow; the infusion of saffron will, perhaps, be found more convenient than any other for performing a few experiments of this kind.

Cloth is dyed black, with the same materials that are used in the preparation of writing ink. Steep a piece of cotton cloth in an infusion of galls or decoction of logwood; dry it and then pass it through a solution of the sulphate of iron. It will be dyed black, the tint deepening on exposure to the air.

By dyeing cloth first with one color and then with another, or with mixtures of different kinds of coloring matter in various proportions, it may be obtained of any particular shade that may be required.

Dip a piece of cloth, dyed blue, (by indigo,) in an infusion of saffron; it will immediately become green, and retain this color permanently. Exp.

Put a number of stuffs, dyed with different kinds of vegetable coloring matter, into an aqueous solution of chlorine, or of the chloride of lime; the color will soon disappear, Exp. but the adjective colors which have been put on without any basis, are destroyed before any of the rest.

#### SECTION IV.—LUTES AND CEMENTS.

A great variety of lutes and cements is required for different chemical operations, and as most of those that are peculiar to particular processes have been already described, in connexion with the method of applying them, it will be sufficient in this place to state those that are more generally useful. They are employed, principally, with the view of rendering joinings tight, so as to confine liquids or gases, or to give support to and protect from the action of the fire, vessels that are apt to be destroyed by exposure to a high temperature.

In a great number of operations, joints may be made sufficiently air-tight by a paste made of linseed meal, pease meal, flour, or starch, (using hot water with the latter,) covering them occasionally where it may be necessary, with a small piece of linen cloth dipped in prepared glue, which may be rendered still more secure by tying a stout thread round them. Small slips of bladder or gum paper will often be found sufficient for all that is required. Powdered gum is frequently used, and a strong solution of it in water should always be kept, which may be thickened with finely pounded chalk, to form the *chalk lute* that has already been so frequently referred to, and which does extremely well for connecting all kinds of glass or earthen tubes that are not to be exposed to a high temperature, or to the action of corrosive liquids or vapors.

Another class of lutes, used for rendering the joinings of apparatus for experiments with the gases; air-tight, consists principally of wax, with various quantities of olive oil, resin, oil of turpentine, or other substances of a similar nature. The *gas lute*, referred to in this work, is composed of one part of wax and three of lard, heating them together till a fluid of a uniform consistence is obtained.

By increasing or diminishing the proportion of wax, it may be easily rendered of various degrees of consistence, so as to make it more fit for a number of different purposes.



For fixing glass tubes into retorts or bottles that are not to be exposed to the action of corrosive liquids, common sealing wax may be employed, tying the glass tube round with thread, till it fits the aperture in which it is to be fixed. A better cement for this purpose, however, is obtained by melting beeswax with an eighth part of common turpentine, which may be made into sticks like sealing wax, melting it and spreading it with a hot iron when required for use. It is less brittle than common sealing wax, and by using a larger portion of turpentine, a soft cement may be obtained, very useful for closing bottles accurately, that contain substances which must not be exposed to the action of the air.

Four parts of resin melted with one of beeswax and then mixed with one part of brick-dust, give a cement that is much employed in joining pieces of apparatus, that are to be permanently fixed together.

Varnishes may often be employed with advantage for rendering a number of joinings tight, as in fastening a bladder or oiled silk bag to a stop-cock, dipping a piece of stout linen cloth in them, and tying it with thread after rolling it round. Thick copal varnish, made by dissolving copal in oil of turpentine, will be found better adapted to this end, than most other kinds of varnishes; it should be purchased from the painters, as the student will find considerable difficulty in preparing it.

To confine acrid vapors, glaziers' putty may be employed; it is made by beating up chalk with drying linseed oil, and is similar in its qualities to the *fat lute*, as it is termed, which is made, by treating clay in the same manner, after drying it thoroughly and reducing it to a fine powder.

A great number of useful lutes and cements may be made, by mixing lime with mucilaginous, albuminous, and gelatinous liquids. Slaked lime in fine powder, made into a paste with the white of an egg, after beating it in a cup, or with a strong solution of glue, sets very quickly, and may be applied most conveniently by spreading it on slips of cloth.

Plaster of Paris is used in the manner described in another place, and by mixing it with a mucilaginous liquid, or a solution of gelatine instead of water, it may be kept for a longer time before it begins to set.

For coating glass vessels so as to enable them to bear a red heat, a mixture of dry sand, powdered clay, and cut thread is to be made into a stiff paste with water, using no more clay than is necessary to make the sand adhere, and mixing the cut thread with the other materials before adding any water; the paste is then to be put equally round the retort and allowed to dry slowly



before it is used. Thin iron wire, wrapped round the coating, serves to bind it together when it becomes hot.

If an iron or earthen vessel is to be coated before exposing it to a high temperature, the directions already given for coating the iron retort, used for the preparation of potassium, present the best method of proceeding in cases of this kind. Most earthen vessels become very porous, however, at a high temperature, and in particular processes, a great portion of the products that would otherwise be obtained, are lost. To obviate this difficulty, Mr Willis succeeded in preparing a lute which is well known by his name; the following is the method of applying it. An ounce of borax is to be dissolved in half a pint of water, and slaked lime added to the solution till a thin paste is obtained. This is to be spread over the retort with a brush, and covered when dry, with a lute made of linseed oil and slaked lime. It may then be put aside for a day or two to dry slowly, when it will be fit for use. The first coating fuses at a high temperature, and forms a glazing over the earthen retort, that prevents vapors from passing through it during distillation; and the second protects it from the fuel, and renders it less liable to be broken as it cools.

#### SECTION V.—ELECTRICITY AND GALVANISM.

An electrophorus, or small electrical machine, will be quite sufficient for performing the experiments described in this work, in which it is necessary to transmit an electric spark through an inflammable gaseous mixture. It is seldom, that the electrical machine is required for any other chemical purpose by the student.

The electrophorus consists of a resinous cake (prepared by melting together equal weights of common resin, shellac, and Venice turpentine, and pouring the compound while still hot, upon a metallic plate or marble slab,) and of a conductor in which the electricity is collected. The resinous cake should be placed upon a metallic plate, and struck lightly with a piece of folded warm flannel, (nothing does better than a fox's tail,) when required for use. The conductor, which is usually made of a piece of thin wood, covered with tin-foil, and provided with a glass handle, is to be placed upon the resinous cake, touched with the finger while in this position, and removed by the glass handle; it will now be charged with electricity, and on bringing any substance near it, a spark will be given off, which may be transmitted through the gaseous mixture to be subjected to its action. Many inflammable gases, mixed with oxygen in the proper proportion, may be analysed in this man-

ner, the quantity of the resulting products indicating their composition. For this purpose, Volta's or Ure's eudiometer is generally employed.

Volta's eudiometer (Fig. 7th, Plate 2d) is made of a stout glass tube, with two platina wires fixed into it near the top, where it is closed, and approaching till they are at a very little distance from each other within it. The mixture to be detonated, is then introduced into the eudiometer over water or mercury, as may be necessary, keeping it in an upright position and wiping the upper part with a cloth, to remove any adhering water or mercury; on connecting one of the wires with the ground by a chain, or touching it with the finger, and bringing the tube near the conductor after it has been properly charged, an electric spark will be immediately transmitted through the mixture, passing between the terminations of the wires within the tube and causing it to explode.

If it be required to examine any of the resulting products, the tube should not be filled more than a third full, and depressed in the liquid to prevent any of the contents being thrown out.

Dr Ure's eudiometer is more convenient than Volta's, for detonating small quantities of inflammable mixtures by means of the electric spark. It consists of a bent tube, closed at one end, with wires fixed into it in the usual manner; the mixture is introduced after filling it with mercury, after which the greater portion of the liquid still remaining in the open extremity of the tube is poured out.

One or both of the wires should terminate in a brass ball, without the eudiometer, that a larger spark may be drawn from the conductor than can be obtained with the wire alone.

Close the open end with the thumb, which is to be placed so as to touch one of the balls; charge the conductor and apply it to the other end. The gaseous mixture at the top expands as the spark passes through it, and presses upon the mercury below, forcing a portion of it into the other extremity of the tube, where the air acts as a spring, and moderates the violence of the explosion. It is necessary to leave as much mercury in the open extremity, as will completely fill the other, should a complete condensation accompany the detonation.

Should the spark not be sufficiently strong to inflame the mixture, a larger quantity of electricity must be collected in a leaden jar by charging the conductor repeatedly, and bringing it in contact with the ball at the top, discharging it in the usual manner.

When it is required to pass a strong electric spark through

any substance, a discharger constructed for the purpose should be employed. This may be made of various sizes and forms, by fixing two pieces of brass rod to the rings of two retort stands, with iron wire; one end of a chain is connected with one of the retort stands, the other end being placed below the Leyden jar after it has been charged; the substance through which the electricity is to be passed, is then placed on a piece of wood or an earthen dish, and the points of the brass rods brought into contact on either side. Another chain is then to be fixed to the second stand, attaching one end of it to a common electrical discharger, and on completing the circuit by bringing it in contact with the ball of the Leyden jar, the electricity immediately passes through the substance placed between the brass wires.

By proceeding in this manner, a number of substances may be easily inflamed or decomposed. Thus, if some powdered resin be thrown on cotton, and placed in the position we have just described, it will immediately take fire on completing the circuit; and fulminating silver and mercury may be exploded in the same manner.

### *Galvanic Battery.*

The galvanic battery is much more frequently employed than the electrical machine, for chemical purposes, and though many improvements have been made in the method of constructing apparatus of this kind, perhaps there is no modification of it that will be found more convenient for performing a few experiments, than the common galvanic trough, in which the plates are arranged in a horizontal position, in the manner proposed by Mr Cruickshanks.\* There are few experiments which cannot be performed with 120 plates, each about four or six inches square, charged with acid liquor of different strengths, according to the purpose for which it may be required. For ordinary experiments, a mixture of two parts of sulphuric acid, one of nitric acid, and eighty of water may be employed; but when it is necessary to produce intense heat or light, one part of sulphuric acid may be mixed with three of nitric acid, and fifteen or twenty of water.

When a single pair of plates is employed, the copper is always the positive pole, (if an acid liquor be used,) the electric current flowing from the zinc to the copper through the liquid, and from the copper to the zinc, when the wires between the two plates are connected together.

In compound galvanic arrangements, the extremity of the

\* See Fig. 6th, Plate 2d.

battery terminating with the plate of zinc, becomes the positive pole; here, however, it must not be forgotten, that this depends solely upon the copper plate to which it is fixed; for, the electric fluid moves from the zinc through the liquid to the copper, as in the preceding case, and the zinc and the copper plates at the extremity of the battery, do not contribute in any degree to the action, and might be removed altogether, without in any degree affecting it; still, however, it is useful to have them attached to the other plates, for the purpose of connecting different troughs together. No difficulty can be experienced by the student, in understanding why the zinc extremity of the battery should be the positive pole in a compound galvanic arrangement; if he looks on a galvanic trough of the usual construction, he must perceive that it is not by the mutual action of the plates that are soldered together and the liquid on either side, that the galvanic electricity is produced, but that it is developed by the action of the liquid on the plates between which it is interposed.

When several troughs are to be connected together, great care must be taken to place them in such a position, that the plates shall all be situated in the same manner with respect to each other, as if they had been placed in a single trough; the zinc extremity of the one being connected with the copper extremity of another.

To detonate gunpowder, fulminating mercury, or any other compound of a similar nature, the best method of proceeding is to put it on a metallic plate connected by a wire with one of the poles of a battery, touching the upper portion with a similar wire connected with the other extremity.

When leaves of gold, silver, or copper, are to be inflamed, a thick brass rod, bent at right angles, is used to suspend them, fixing it at the positive or negative pole of the battery, and touching the metallic leaf with a plate of iron or zinc connected with the other pole of the battery, gilded previously, to reflect the light better.

When the galvanic electricity is transmitted through charcoal, small pencils, about two or three inches in length, and one-tenth of an inch in thickness, may be employed; they are prepared most conveniently by heating pieces of willow, cut to a proper size, in a covered crucible filled with powdered charcoal.

A few experiments, illustrating the power of galvanism in effecting decomposition, may be easily performed.

Fill a tube with a diluted solution of sulphate of soda, and invert it under water, fixing it into the central aperture of a glass vessel, which has an additional opening on either side of the centre, (a Woulfe's bottle would answer);



then put two gold or platina wires, (connected with the positive and negative poles of the battery by copper wires,) into the tubulures or openings at the side, through corks, and immediately a stream of gas will rise from both wires, and collect in the tube; the wires must not close the apertures in the corks completely, but allow part of the liquid to pass out, as it is displaced by the gas. On examining the mixture, it will be found to be composed of one measure of hydrogen, and half a measure of oxygen, (a portion of water being decomposed,) which may be easily detonated by withdrawing the tube and applying a light. Oxygen is evolved at the positive, and hydrogen at the negative wire, and they may be collected in separate vessels, by a modification of the apparatus, represented in the above figure.

Take two glasses, filled with the same diluted solution of the sulphate of soda, connecting them together by a piece of moistened cloth, and putting a wire attached to the positive pole of the battery into one of the glasses, and one connected with the negative pole into the other. Water will be decomposed, as in the preceding experiment, and also the sulphate of soda; in the former case, however, the sulphuric acid and soda detached from one another, meeting again in the same vessel, immediately combine and form sulphate of soda, so that they cannot be obtained separate from one another; but, in the present case, all the acid passes into the glass connected with the positive pole, and the alkali into the other. If the solution be tinged with a little of the blue infusion of cabbage, the usual changes of color will take place, and indicate the change that is going on. Exp.

By changing the position of the wires, the acid may be made to pass into the glass which previously contained the alkali in excess, while the latter takes the place of the acid.

In both these experiments, the copper wire connected with the positive pole of the battery, must have a piece of gold or silver wire attached to it, which alone is to be introduced into the liquid; a copper wire being speedily oxydated as the water is decomposed, and the oxyde combining at the same time with the acid of the sulphate, which is separated at the same pole.

#### SECTION VI.—ACIDIMETRY AND ALKALIMETRY.

The method of ascertaining the exact quantity of free acid or alkali in a given weight of any solution or mixture of solid substances, is a subject of great importance, as it not only enables



us to determine their value in a commercial point of view, but also to adjust the quantities of materials required for different processes and experiments. The strength, indeed, of every acid and alkaline solution used for chemical experiments should be known, and it will be found to save a great deal of time, if the solutions that are employed, are always prepared of the same strength.

To ascertain the strength of any acid, (an operation that is now termed *acidimetry*), all that is necessary, is to take a given weight of an alkali, and see what quantity it can neutralise. Thus, if we wish to ascertain how much real sulphuric acid there is in diluted sulphuric acid, we have only to take a given quantity of an alkali, or alkaline carbonate, that of potash for example, and ascertain what weight of the acid liquid is required to neutralise it; and as we know that 40 parts of sulphuric acid neutralise exactly 48 of potash, we must allow 40 parts of acid for every 48 of potash, neutralised. If carbonate of potash be employed, then every 70 parts that are neutralised indicate the presence of 40 of acid, containing exactly 48 of potash.

For example; suppose we dissolve 70 grains of the carbonate of potash in water, and find that it requires exactly 160 grains of the diluted sulphuric acid to neutralise it; then these 160 grains must have contained exactly 40 of dry sulphuric acid, or 25 per cent.; consequently, in all experiments where a given weight of acid is required, we must take four times its weight of this solution, to have the proper quantity.

In the same manner, we may easily ascertain the quantity of real acid in most other strong or diluted acid liquids; 48 parts of real potash (or 70 of the carbonate) indicating the presence of 54 of nitric acid, 37 of muriatic, and quantities of other acids represented by their respective equivalents, when they are capable of completely neutralising potash.

It must be obvious, that, on the same principle, we may ascertain the quantity of free alkali in any alkaline solid or liquid; and the process for this purpose has now received the name of *alkalimetry*. Thus, if we should wish to ascertain the quantity of soda in a specimen of kelp, the value of which depends principally on the quantity of this ingredient which it contains, we have first to extract all the soluble matter from a given weight of it. This is done by digesting it in water and washing the residuum repeatedly with water, till no more soluble matter is taken up. Then we neutralise the mixed liquids with a diluted acid of a given strength, and thus easily calculate the quantity of free soda which the kelp contains. For example; if the sol-

luble matter extracted from 500 grains of kelp should require 80 grains of diluted sulphuric acid, containing 25 per cent. of real acid, to neutralise it, then they must have contained sixteen grains of soda; for 80 parts of the diluted acid contain 20 of real acid, and 40 of this acid neutralise 32 of soda; but if 40 of sulphuric acid neutralise 32 of soda, then half the quantity will neutralise sixteen of this alkali.

On the same principle, the quantity of potash in pearlash and potashes may be ascertained, and of real ammonia in the water of ammonia, substituting the respective equivalents of these substances for that of soda.

For ascertaining when the neutralisation is completed in these processes, nothing will be found more convenient than the infusion of blue cabbage, as a very slight excess of acid or alkali changes its color to red or green. Test-papers, prepared in the manner already described, may be used for the same purpose.

A pure carbonate of potash may be easily prepared for these experiments, by fusing the crystallised bi-carbonate in a platina crucible. The best method of proceeding, is to pour it out on a clean iron plate, after it has been kept in a state of fusion for five or ten minutes in a platina crucible, putting a fragment of it into a small cup of water accurately equipoised in a balance, and ascertaining its weight by again counterbalancing the cup. The object of proceeding in this manner is to prevent it gaining any increase of weight by attracting water from the air, and both the cup and the water having been previously balanced, the additional weights required to bring the scales again to an equilibrium represent the exact weight of the fragment put into the water.

Instead of using the pure carbonate, the bi-carbonate itself is occasionally employed for the same purpose, and as every 101 parts contain 48 of potash, we must allow 40 parts of sulphuric acid for every 101 parts of this salt which are neutralised.

Crystallised carbonate of soda is occasionally used for the same purpose, but the quantity of water of crystallisation in it is said to vary, sometimes containing ten and sometimes eleven equivalents of water, combined with one of the dry carbonate.

With respect to the method of measuring the quantity of the acid liquid employed in ascertaining the quantity of free alkali in any solution, there are two methods of proceeding. According to one, a liquid measure is used, divided into tenths of a cubic inch, the diluted acid being made of such a strength, that every tenth of a cubic inch shall contain a sufficient quantity of real acid to neutralise half a grain or a grain of potash or soda.

The other method consists in adding a given weight of the

acid to the alkaline solution, and calculating the quantities in the manner described. This is the method I am inclined to prefer, and those who adopt this plan, will find it convenient to proceed in the following manner. Take a glass bottle of a conical form, flat at the bottom and not very tall, that it may rest steadily upon one of the scales of a balance. Put a pipette with a very fine point and some of the test-acid with which the strength of the alkaline solution is to be ascertained, into it, and counterpoise this scale with weights, or shot and sand. Then fill the ball of the pipette with the diluted acid by sucking out part of the air, remove it carefully, and hold it over the solution to be neutralised in such a manner that the acid liquid shall drop out slowly; the smallest portion of acid can be added in this manner with facility, and on replacing the pipette after the neutralisation of the liquid, the weights necessary to restore the former equilibrium indicate the quantity of diluted acid employed, from which the strength of the alkaline solution may be calculated as before.

SECTION VII.—TABLES OF WEIGHTS AND MEASURES, OF THE CORRESPONDENCE BETWEEN FAHRENHEIT'S, REAUMUR'S, AND THE CENTIGRADE THERMOMETERS, AND OF FREEZING MIXTURES.

*Weights and Measures.*

WEIGHTS.

There are three sorts of weights in use in Great Britain and the United States, viz. the *Troy* weight, *Avoirdupois* weight, and *Apothecaries'* weight. The last of these is usually employed by American Pharmacutists.

TROY WEIGHT.

24 grains	=	1 pennyweight.
20 pennyweights	=	1 ounce.
12 ounces	=	1 pound.

or,

Grains.		Pennyweights.		Ounces.		Pound.
24	=	1	=	$\frac{1}{20}$	=	$\frac{1}{240}$
480	=	20	=	1	=	$\frac{1}{12}$
5760	=	240	=	12	=	1

The pound Avoirdupois contains 7000 grains, each of which is equal to a Troy grain, being thus heavier than the Troy pound by 1240 grains.

## AVOIRDUPOIS WEIGHT.

16 drachms	=	1 ounce	=	437.5 Troy grains.
16 ounces	=	1 pound	=	7000
28 pounds	=	1 quarter	=	196000
4 quarters	=	1 cwt. or 112 lbs.	=	784000
20 cwt.	=	1 ton	=	15680000

Pounds.	Ounces.	Drachms.	Troy weight.
1	= 16	= 256	= 7000 grains.
$\frac{1}{16}$	= 1	= 16	= 437.5
$\frac{1}{256}$	= $\frac{1}{16}$	= 1	= 27.34375

The pound in Apothecaries' weight is equal to the Troy pound, containing 5760 grains, but is differently subdivided.

## APOTHECARIES' WEIGHT.

1 pound lb.	=	12 ounces	=	5760 Troy grains.
1 ounce ℥	=	8 drachms	=	480
1 drachm ℥	=	3 scruples	=	60
1 scruple ℥	=	20 grains	=	20

Pounds.	Ounces.	Drachms.	Scruples.	Grains.
1	= 12	= 96	= 288	= 5760
	= 1	= 8	= 24	= 480
		= 1	= 3	= 60
			= 1	= 20

The following tables show the correspondence between the Troy, Avoirdupois, and Apothecaries' weights.

Troy weight.	Avoirdupois:			Apothecaries'.
	oz.	dr.	grains.	
1 pound	= 13	2	17.8125	= 1 pound.
1 ounce	= 1	1	15.1562	= 1 ounce.
1 pennyweight	= 0	0	24	= 1 scruple, 4 grains.

Avoirdupois.	Troy weight.			Apothecaries'.			
	lb.	oz.	dwt.	lb.	oz.	dr.	sc.
1 pound	= 1 lb.	2 oz.	11 dwt.	16 gr.	= 1 lb.	2 oz.	4 dr. 2 sc.
1 ounce	= 0	0	18	5.5	= 0	0	7 0 17.5 gr.
1 drachm	= 0	0	1	3.34	= 0	0	0 1 7.34

Apothecaries'.	Troy weight.			Avoirdupois.	
	lb.	oz.	dwt.	lb.	oz.
1 pound	= 1 pound	= 13 oz.	2 dr.	17.8125 gr.	
1 ounce	= 1 ounce	= 1	1	15.1562	
1 drachm	= 2 dwt.	12 gr.	= 0	2	5.3125
1 scruple	= 0	20	= 0	0	20

FRENCH DECIMAL MEASURE OF WEIGHT.—GRAMME = 15.4410  
TROY GRAINS.

Milligramme	=	.0154				IMPERIAL, OR	
Centigramme	=	.1544				TROY WEIGHT.	
Decigramme	=	1.5444	Lbs.	Ozs.	Dwts.	Grains.	
Gramme	=	15.4440	= 0	0	0	15.4	
Decagramme	=	154.4402	= 0	0	6	10.44	
Hecagramme	=	1544.4023	= 0	3	4	8.40	
Kilogramme	=	15444.0234	= 2	8	3	12	
Myriagramme	=	154440.2344	= 26	9	15	0	

### MEASURES.

The Imperial Standard Gallon contains ten pounds Avoirdupois weight of distilled water, weighed in air at 62° Fah. and 30° Barom., or 277.274 cubic inches; or 12 lbs., 1 ounce, 16 pennyweights, and 16 grains Troy, = 70,000 grains of distilled water.

#### IMPERIAL MEASURE.

1 quarter	=	8 bushels.
1 bushel	=	4 pecks.
1 peck	=	2 gallons.
1 gallon	=	4 quarts.
1 quart	=	2 pints.

or,

Troy grs.	Avoird. lb.	Cub. inch.	Pint.	Quar.	Galls.	Pecks.	Bush.
8750 =	1.25 =	34.659 =	1				
17500 =	2.5 =	69.318 =	2 =	1			
70000 =	10 =	277.274 =	8 =	4 =	1		
	20 =	554.548 =	16 =	8 =	2 =	1	
	80 =	2218.192 =	64 =	32 =	8 =	4 =	1
	640 =	17745.536 =	512 =	256 =	64 =	32 =	8 = 1

The gallon of the former wine measure and of the present Apothecaries' measure contains 58.413 Troy grains weight of distilled water, or 231 cubic inches, the ratio being nearly as 6 to 5, or as 1 to 0.8331.

#### APOTHECARIES' MEASURE.

1 gallon	=	8 pints.
1 pint	=	16 ounces.
1 ounce	=	8 drachms.
1 drachm	=	60 drops, or minims.

or,

Gallon.	Pints.	Ounces.	Drachms.	Minims.	Grains Troy.	Cub. Inches.
1 =	8 =	128 =	1024 =	61440 =	58463 =	231
	1 =	16 =	128 =	7680 =	7292.3 =	28.8
		1 =	8 =	480 =	455.6 =	1.8
			1 =	60 =	56.9 =	0.2



French Decimal Measure of Capacity, Litre = 61.028 British cubic inches, or 2.113 Apothecaries' pint, or 1.7608 Imperial pint, or 31.104 Troy ounces of water.

Millilitre =	.0610	IMPERIAL MEASURE.				
Centilitre =	.6102					
Decilitre =	6.1028	Pints.	Quarts.	Gallons.	Pecks.	Bushels.
Litre =	61.0280=	1.7608=0		0	0	0
Decalitre =	610.2800=	17.608=	8.804=	2.201=	1.100	
Hecatolitre=	6102.8000=	176.08=	88.04=	22.01=	11.005	
Kilolitre =	61028.0000=	1760.8=	880.4=	220.1=	110.05=	27.51
Myrialitre =	610280.0000=	17608=	8804=	2201=	1100.5=	275.1

TABLE SHOWING THE WEIGHT IN GRAINS OF VARIOUS MEASURES OF DIFFERENT FLUIDS.

	Spr. gr.	1 Pint.	1 Ounce.	1 Drachm.	1 Minim.
Distilled water	1000	7305	456.5	57	.9
Sulphuric ether	720	5259	328.6	41.07	.68
Alcohol	796	5814	363.4	45.0	.78
Solution of ammonia	926	6825	425	53.2	0.84
Muriatic acid	1280	9450	590.6	73.8	1.23
Nitric acid	1500	10957	654	85.5	1.35
Sulphuric acid	1845	13477	840	105	1.65

TABLES SHOWING THE CORRESPONDENCE BETWEEN FAHRENHEIT'S, REAUMUR'S, AND THE CENTIGRADE THERMOMETERS.

The space between the boiling and freezing points of water is divided into

180	parts or degrees	in Fahrenheit's Thermometer.
100	"	in the Centigrade.
80	"	in Reaumur's.

Therefore,

Fahren.	Cent.	Reaum.
180 degrees	= 100	= 80
1	= $\frac{5}{9}$	= $\frac{4}{5}$
$1\frac{4}{5}$	= 1	= $\frac{4}{5}$
$2\frac{4}{5}$	= $1\frac{1}{5}$	= 1

In Fahrenheit's thermometer the graduation begins at 32 degrees below the freezing point of water. The Centigrade and Reaumur's commence at this point.

Accordingly,

1. To reduce Centigrade degrees to Fahrenheit's, multiply by 9, divide by 5, and add 32. Thus,

$$40 \text{ C.} \times 9 = 360; \frac{360}{5} = 72; 72 + 32 = 104 \text{ Fahrenheit.}$$

$$\text{Formula } \frac{\text{C.} \times 9}{5} + 32 = \text{F.}$$

2. To reduce Fahrenheit's to Centigrade's, subtract 32, multiply by 5, and divide by 9. Thus,

$$104 \text{ Fah.} - 32 = 72; 72 \times 5 = 360; \frac{360}{9} = 40 \text{ Centigrade.}$$

$$\text{Formula } \frac{\text{F.} - 32 \times 5}{9} = \text{C.}$$

3. To reduce Reaumur's to Fahrenheit's, multiply by 9, divide by 4, and add 32. Thus,

$$R. 32 \times 9 = 288; \frac{288}{4} = 72; 72 + 32 = 104 \text{ Fah.}$$

$$\text{Formula } \frac{R. \times 9}{4} + 32 = F.$$

4. To reduce Fahrenheit's to Reaumur's, subtract 32, multiply by 4, and divide by 9. Thus,

$$\text{Fah. } 104 - 32 = 72; 72 \times 4 = 288; \frac{288}{9} = 32 \text{ R.}$$

$$\text{Formula } \frac{F. - 32 \times 4}{9} = R.$$

5. To reduce Reaumur's to the Centigrade, multiply by 5, and divide by 4. Thus,

$$R. 32 \times 5 = 160; \frac{160}{4} = 40 \text{ Centigrade.}$$

$$\text{Formula } \frac{R. \times 5}{4} = C.$$

6. To reduce the Centigrade to Reaumur's, multiply by 4, and divide by 5. Thus,

$$C. 40 \times 4 = 160; \frac{160}{5} = 32 \text{ Reaum.}$$

$$\text{Formula } \frac{C. \times 4}{5} = R.$$

A TABLE SHOWING THE CORRESPONDENCE BETWEEN FAHRENHEIT'S, REAUMUR'S, AND THE CENTIGRADE THERMOMETERS.

F.	C.	R.	F.	C.	R.	F.	C.	R.	F.	C.	R.	F.	C.	R.
-40	-40	-32	-14	-10	-8	68	20	16	122	50	40	176	80	64
-31	-35	-28	-23	-5	-4	77	25	20	131	55	44	185	85	68
-22	-30	-24	-32	0	0	86	30	24	140	60	48	194	90	72
-13	-25	-20	-41	5	4	95	35	28	149	65	52	203	95	76
-4	-20	-16	-50	10	8	104	40	32	158	70	56	212	100	80
+5	-15	-12	-59	15	12	113	45	36	167	75	60	&c.		

It will be observed, that in this table the numbers in Fahrenheit's scale advance by 9, in the Centigrade by 5, and in Reaumur's by 4.

Wedgewood's pyrometer begins at 1077 Fahrenheit; and each degree is equal to 130 of Fahrenheit. Therefore,

$W. \times 130 + 1077 = \text{Fah.}$ ; or according to Morveau, who affirms that the degrees in Wedgewood's pyrometer are estimated too high;  $W. \times 62.5 + 517.579 = \text{Fah.}$ ; and also  $\text{Fah.} - 1077 \div 130 = W.$ ; or according to Morveau,  $\text{Fahren.} - 517.579 \div 62.5 = W.$  It is an instrument, however, that is scarcely ever used now, as the indications which it gives cannot be depended upon.

## TABLE

OF

## FRIGORIFIC MIXTURES WITHOUT ICE OR SNOW.

Mixtures. Quantity of Materials by Weight.	Reduction of Temperature.	Number of degrees the thermome- ter falls.
Diluted sulphuric acid,* 4 parts. Sulphate of soda,† 5	From + 50 to + 3.	47
Muriatic acid, 5 Sulphate of soda, 8	From + 50 to 0.	50
Diluted nitric acid,‡ 4 Phosphate of soda, 9	From + 50 to — 12.	62
Water, 1 Nitrate of ammonia, 1	From + 50 to + 4.	46
Water, 16 Nitrate of potash, 5 Muriate of ammonia, 5	From + 50 to + 10.	40
Diluted nitric acid, 4 Nitrate of potash, 2 Muriate of ammonia, 4 Sulphate of soda, 6	From + 50 to — 10.	60

\* Composed of equal weights of strong sulphuric acid and water, allowed to cool before using it.

† Crystallised sulphate of soda, which should be reduced to powder before it is used.

‡ Prepared by mixing one part of water and two of the strong acid.

## TABLE

OF

## FRIGORIFIC MIXTURES WITH ICE OR SNOW.

Mixtures. Quantity of Materials by Weight.	Reduction of Temperature.	Number of degrees the thermome- ter sinks.
Snow,* 3 parts. Diluted sulphuric acid, 2	From + 32 to — 23.	55
Snow, 7 Diluted nitric acid, 4	From + 32 to — 30.	62
Snow, 8 Muriatic acid, 5	From + 22 to — 27.	59
Snow, 2 Chloride of sodium, 1  Snow,* 5 Chloride of sodium, 2 Muriate of ammonia, 1	<div> <div>From any Temperature</div> <div> <div>to — 5.†</div> <div>to — 12.</div> </div> </div>	
Snow, 2 Crystallised muriate of lime, 3	From + 32 to — 50.	82
Snow, 1 Crystallised muriate of lime, 3	From — 40 to — 73.	33
Snow, 8 Diluted sulphuric acid, 10	From — 68 to — 91.	23

\* When snow cannot be procured, pounded ice may be substituted for it.

† The temperature of a mixture of common salt and snow has fallen to — 9, when a very large quantity of materials was mixed together.

# INDEX.

	Page		Page
Acetated cerusse, . . . . .	394	Acid, hydrosulphuric, . . . . .	116
Acetates, . . . . .	309, 368	hyponitrous, . . . . .	88
Acetous fermentation, . . . . .	191	hyperoxymuriatic, . . . . .	247
Acidifying principle, . . . . .	43	hypophosphorous, . . . . .	133
Acid, definition of, . . . . .	88	hyposulphuric, . . . . .	115
Acid, acetic, . . . . .	192	hyposulphurous, . . . . .	115
acetous, . . . . .	192	iodic, . . . . .	261
antimonious, . . . . .	416	kinic, . . . . .	223
antimonic, . . . . .	417	lampic, . . . . .	170
arsenic, . . . . .	439	malic, . . . . .	203
arsenious, . . . . .	430	manganesic, . . . . .	450
auric, . . . . .	501	manganeseous, . . . . .	450
benzoic, . . . . .	201	margaric, . . . . .	228
boracic, . . . . .	235	meconic, . . . . .	217
carbonic, . . . . .	147	molybdic, . . . . .	464
chloric, . . . . .	247	mucic, . . . . .	205
chloriodic, . . . . .	266	muriatic, . . . . .	248
chlorocyanic, . . . . .	256	nitric, . . . . .	90
chlorocarbonic, . . . . .	255	nitromuriatic, . . . . .	97
chlorochromic, . . . . .	459	nitrous, . . . . .	89
chromic, . . . . .	458	of sea salt, . . . . .	248
citric, . . . . .	197	oleic, . . . . .	224
columbic, . . . . .	466	oxalic, . . . . .	198
cyanic, . . . . .	179	oxychloric, . . . . .	248
ellagic, . . . . .	202	oxymuriatic, . . . . .	235
ferrocyanic, . . . . .	186	perchloric, . . . . .	243
ferruretted chyazic, . . . . .	188	phosphatic, . . . . .	135
fluoboric, . . . . .	270	phosphorous, . . . . .	132
fluoboracic, . . . . .	270	phosphoric, . . . . .	133
fluochromic, . . . . .	459	prussic, . . . . .	179
fluoric, . . . . .	269	pyroligneous, . . . . .	191
fluosilicic, . . . . .	271	pyrotartaric, . . . . .	195
fulminic, . . . . .	179, 492	saccholactic, . . . . .	205
gallic, . . . . .	202	selenic, . . . . .	125
hydriodic, . . . . .	262	silicic, . . . . .	371
hydrobromic, . . . . .	267	silicofluoric, . . . . .	271
hydrochloric, . . . . .	248	sorbic, . . . . .	203
hydrocyanic, . . . . .	179	stibic, . . . . .	417
hydrofluoric, . . . . .	269	stibious, . . . . .	417
hydroxanthic, . . . . .	232	succinic, . . . . .	203
hydroselenic, . . . . .	125	sulphuric, . . . . .	109



	Page		Page
Acid, sulphurous, . . . . .	105	Ammonia, supposed metallic base	
sulphuretted chyazic, . . . . .	188	of, . . . . .	330
sulphocyanic, . . . . .	188	acetate of, . . . . .	338
tartaric, . . . . .	194	muriate of, . . . . .	336, 337
titanic, . . . . .	467	nitrate of, . . . . .	338
tungstic, . . . . .	463	benzoate, . . . . .	340
uric, . . . . .	514	carbonate, . . . . .	334
vitriolic, . . . . .	109	hydriodate, . . . . .	333
Acidimetry, . . . . .	535	phosphate, . . . . .	339
Acids, vegetable, table of, . . . . .	190	sulphate, . . . . .	336
Acidifying principle, . . . . .	43	succinate, . . . . .	340
Acidulous waters, . . . . .	520	water of, . . . . .	331
Adipocere, . . . . .	510	liquid, . . . . .	331
Adjective colors, . . . . .	528	Ammonium, . . . . .	320
Ærated magnesia water, . . . . .	361	Ammoniuret of copper, . . . . .	402
Ærugo æris, . . . . .	405	Anhydrite, . . . . .	348
Affinity, . . . . .	15, 17, 18	Anhydrous, . . . . .	45, 109
Air, atmospheric, . . . . .	99	Animal substances, . . . . .	510
Alabaster, . . . . .	348	Animal and vegetable analysis, . . . . .	517
Albumen, . . . . .	511	Antimonial wine, . . . . .	416, 427
Alcoates, . . . . .	166	Antimonites, . . . . .	416
Alcohol, . . . . .	160	Antimony, . . . . .	411
how formed, . . . . .	161, 162	butter of, . . . . .	419
absolute, . . . . .	163, 164	crude, . . . . .	413
ammoniatum, . . . . .	322	chloride of, . . . . .	419
Lowitz's table of, . . . . .	165	crocus of, . . . . .	415
poisonous quality, . . . . .	167	glass of, . . . . .	415
quantity of, in liquors, . . . . .	165	sulphuret of, . . . . .	412, 420
uses of, . . . . .	166	golden sulphuret, . . . . .	422
unnecessary, . . . . .	166, 167	muriate of, . . . . .	419, 437
Algaroth's powder, . . . . .	415	oxyde of, with phos- phate of lime, . . . . .	418
Alkalimetry, . . . . .	536	and potash, tartrate of, . . . . .	423
Alkali, . . . . .	210	protoxyde of, . . . . .	415
vegetable, . . . . .	210	sulphate of, . . . . .	423
mineral, . . . . .	318	regulus of, . . . . .	413
fixed, . . . . .	318	tests of, . . . . .	425
volatile, . . . . .	331	Aqua, . . . . .	71
Alkalifiable metals, . . . . .	287 to 341	fortis, . . . . .	93
Alkaline waters, . . . . .	521	oxy-muriatica, . . . . .	240
Alkaloids, . . . . .	211	regia, . . . . .	97
Alloys, what? . . . . .	507	Aqueous chlorine, . . . . .	240
Alkaline geofiable metals, . . . . .	341	Apparatus, Nooth's, . . . . .	149
Alum, . . . . .	365	Woulfe's, . . . . .	ix
curd, . . . . .	366	Arbor dianæ, . . . . .	494
whey, . . . . .	366	saturni, . . . . .	389
Alumen exsiccatum, . . . . .	367	Argentiferous flowers of antimony, . . . . .	414
Alumina, . . . . .	363	Aromatic spirit of sulphuric ether, . . . . .	172
acetate of, . . . . .	363	Arsenic, regulus of, . . . . .	428
Aluminium, . . . . .	362	reduction of, . . . . .	429
Amalgams, . . . . .	507	mineral solution of, . . . . .	430
for electrical uses, . . . . .	508	protosulphuret of, . . . . .	440
for looking-glasses, . . . . .	508	yellow sulphuret of, . . . . .	440
Amber, acid from, . . . . .	203		
Ammonia, . . . . .	102, 330		
characters of, . . . . .	330, 332		

	Page		Page
Arsenic, oxyde of, . . . . .	430	Bi-sulphuretted hydrogen, . . . . .	123
modes of testing, 436, 437, . . . . .	438	Bi-tartrate of potash, . . . . .	309
its antidotes, . . . . .	442	Black drop, . . . . .	193
Arsenical experiments, 434, 435 . . . . .	440	Black flux, . . . . .	311
Arseniuretted hydrogen, . . . . .	440	Black lead, . . . . .	377
Atmospheric air, . . . . .	99	crucibles, . . . . .	377
its composition, . . . . .	99	Bleaching salts, . . . . .	350
properties of, . . . . .	101	Blende zinc, . . . . .	407
Atomic theory, . . . . .	9, 10	Blowpipe, alcoholic, . . . . .	166
Attraction of cohesion, . . . . .	4	compound, . . . . .	69
of composition, . . . . .	6	oxy-hydrogen, . . . . .	69
Aurates, . . . . .	501	Blue cabbage, infusion of, . . . . .	525
Azote, . . . . .	76	Blue, Prussian, . . . . .	385
Balance, . . . . .	vii	Blue vitriol, . . . . .	401
hydrostatic, . . . . .		Borax, . . . . .	329
Baldwin's phosphorus, . . . . .	350	calcined, . . . . .	329
Barilla, . . . . .	318	Boron, . . . . .	233
Barium, . . . . .	351	chloride of, . . . . .	256
protoxyde of, . . . . .	352	Borotartrate of potash, . . . . .	310
peroxyde of, . . . . .	353	Brass, . . . . .	507
deutoxyde of, . . . . .	353	Brazil wood, . . . . .	528
sulphuret, . . . . .	353	Bromides, . . . . .	268
Barytes, . . . . .	352	Bromine, . . . . .	266
its distinctive marks, . . . . .	352	cyanide of, . . . . .	268
carbonate, . . . . .	354	Bronze, . . . . .	507
chlorate, . . . . .	355	Brucine, . . . . .	224
muriate, . . . . .	354	Burnt alum, . . . . .	367
nitrate, . . . . .	354		
sulphate, . . . . .	353	Calamine, . . . . .	407
tests of, . . . . .	352	Calcareous nitre, . . . . .	302
Barytum, . . . . .	351	Calcium, . . . . .	341
Basis, . . . . .	528	chloride of, . . . . .	347
Benzoates, . . . . .	201	oxyde of, . . . . .	343
Benzoin, . . . . .	201	phosphuret of, . . . . .	346
Bi-carbonate of potash, . . . . .	308	Calculi, analysis of, . . . . .	513, 516
Bi-carburet of hydrogen, . . . . .	159	biliary, . . . . .	513
nitrogen, . . . . .	175	cystic oxyde, . . . . .	516
Bi-chloride of mercury, . . . . .	480	fibrinous, . . . . .	516
Bi-hydrogenated carbon, . . . . .	156	fusible, . . . . .	516
phosphorus, . . . . .	139	phosphate of lime, . . . . .	515
Biliary calculi, . . . . .	513	phosphate of ammonia . . . . .	515
Binary, . . . . .	45	and magnesia, . . . . .	515
Bi-phosphuretted hydrogen, . . . . .	139	oxalate of lime, . . . . .	514
Bis, . . . . .	45	urate of ammonia, . . . . .	514
Bismuth, . . . . .	446	uric acid, . . . . .	514
nitrate of, . . . . .	448	xanthic oxyde, . . . . .	516
bi-nitrate, . . . . .	448	Calorimeter, . . . . .	36
oxyde, . . . . .	447	Calomel, . . . . .	483
subnitrate, . . . . .	448	Caloric, . . . . .	21
tartrate, . . . . .	449	what is it ? . . . . .	36
Bi-sulphate of potash, . . . . .	306	free, . . . . .	21
Bi-sulphuret of carbon, . . . . .	231	latent, . . . . .	21
how formed, . . . . .	231	radiant, . . . . .	22
characters of, 232, 233 . . . . .		sensible, . . . . .	21
		specific, . . . . .	34

	Page		Page
Camphor, . . . . .	229	Choke damp, . . . . .	151
artificial, . . . . .	230	vapor, . . . . .	151
Capacity for heat, . . . . .	35	Cholesterine, . . . . .	513
Carbon, . . . . .	140	Chrome, . . . . .	456
characters of, . . . . .	140	yellow, . . . . .	461
bi-hydruret, . . . . .	156	Chromate of barytes, . . . . .	460
bi-sulphuret, . . . . .	231	lead, . . . . .	461
hydruret, . . . . .	152	mercury, . . . . .	461
perchloride, . . . . .	255	potash, . . . . .	459
its absorbing power, . . . . .	142	Cinchonia, . . . . .	223
its compounds, . . . . .	140	Cinnabar, . . . . .	488
protochloride, . . . . .	255	Citric acid, . . . . .	197
subchloride, . . . . .	255	Coal gas, . . . . .	153
Carbonates of potash, . . . . .	306	Cobalt, . . . . .	461
soda, . . . . .	325	muriate, . . . . .	462
Carbonic oxyde, . . . . .	144	oxydes, . . . . .	462
characters of, . . . . .	147	Cochineal, . . . . .	527
preparation of, . . . . .	145, 146	Cold, a sedative, . . . . .	41
its discovery, . . . . .	145	Coloring matter, . . . . .	525
Carburetted hydrogen, . . . . .	152	Colors, adjective, . . . . .	528
Carmine, . . . . .	527	substantive, . . . . .	528
Carbosulphurets, . . . . .	232	Columbium, . . . . .	465
Cassius, purple preeipitate of, . . . . .	502	Combustion, . . . . .	271
Caustic, common, . . . . .	294	Copal varnish, . . . . .	530
lunar, . . . . .	497	Copper, and its compounds, . . . . .	397
Cements, . . . . .	530	metallie, not poisonous, . . . . .	398
Cerusse, . . . . .	394	Copperas, . . . . .	378
acetated, . . . . .	394	Corrosive sublimate, . . . . .	480
Chalybeate waters, . . . . .	521	its poisonous character, . . . . .	482
Chameleon mineral, . . . . .	452	Cream of tartar, . . . . .	309
Charecoal, . . . . .	141	of lime, . . . . .	343
how made, . . . . .	141	Crocus of antimony, . . . . .	414
animal, . . . . .	142	martis, . . . . .	375
its indestructibility, . . . . .	143	Crucibles, black lead, . . . . .	377
its chemical relations, . . . . .	144	Crystal, mineral, . . . . .	304
its medical uses, . . . . .	144	Cubic nitre, . . . . .	329
Chauffer, . . . . .	ix	alum, . . . . .	367
Chemical affinity, . . . . .	6	Cupellation, . . . . .	495
equivalents, . . . . .	10	Cuprum ammoniacum, . . . . .	402
Chlorate of potash, . . . . .	313	Cyanide of iodine, . . . . .	266
Chloric ether, . . . . .	255	of potassium, . . . . .	300
Chloridic theory, . . . . .	274	Cyanogen, . . . . .	175
Chlorides, 253, 256, 255, 266, 300, . . . . .	314, 347, 350, 391	preparation of, . . . . .	176
Chlorine, . . . . .	235	characters of, . . . . .	177
characters of, . . . . .	239, 241	its chemical relations, . . . . .	178
cyanide of, . . . . .	256	Cyanurets, . . . . .	178
experiments with, . . . . .	241	Dalby's earminative, . . . . .	361
hydrocarburet of chloro- . . . . .	255	Definite proportions, . . . . .	10
rine, . . . . .	255	Definition of chemistry, . . . . .	16
preparation of, . . . . .	237	Deliquescent, . . . . .	301
peroxyde of, . . . . .	245	Derosne, salt of, . . . . .	218
protoxyde, . . . . .	243	Delphine, . . . . .	226
table of compounds, . . . . .	236	Dephlogisticated marine acid, . . . . .	235
Chlorurets, . . . . .	243	Deutoxyde of nitrogen, . . . . .	84
		Diamond, . . . . .	140

	Page		Page
Di-acetate of lead, . . . . .	396	Galena, . . . . .	392
Di-chloride, . . . . .	420	Gallates, . . . . .	203
Dilatation, . . . . .	25	Galvanism, . . . . .	533, 535
Di-nitrate, . . . . .	393, 448	Gas, exhilarating, . . . . .	83
Dobereiner's lamp, . . . . .	67	Gas lute, . . . . .	529
Dutch gold, . . . . .	507	Gases, to find specific gravity of, . . . . .	49
Effervescing draughts, . . . . .	152	Gelatine, . . . . .	512
Efflorescent, . . . . .	301	Geofiable metals, . . . . .	362
Elaine, . . . . .	227	Gilding, . . . . .	509
Elective affinity, . . . . .	15	Glacial oil of vitriol, . . . . .	109
Electricity, . . . . .	531, 533	Glass, . . . . .	372
Emetic tartar, . . . . .	414	of antimony, . . . . .	415
Emetine, . . . . .	226	Glauber's salts, . . . . .	323
Emulsion, . . . . .	226	effloresced, . . . . .	324
Equilibrium of heat, . . . . .	22	Gliadine, . . . . .	209
Epsom salts, . . . . .	359	Glucine, . . . . .	372
Ether, . . . . .	167	Glucinium, . . . . .	372
nitric, . . . . .	172	Gluten, . . . . .	208
sulphuric, . . . . .	168	Gold coin, . . . . .	508
Ethereal gas, . . . . .	172	Gold and its compounds, . . . . .	500, 503
Ethiops martial, . . . . .	374	Golden calf, . . . . .	445
mineral, . . . . .	487	Golden sulphuret of antimony, . . . . .	422
per se, . . . . .	468	Green vitriol, . . . . .	378
Euchlorine, . . . . .	245	Gum, . . . . .	204
Eudiometers, . . . . .	100, 532	resin, . . . . .	231
Exhilarating gas, . . . . .	83	Gunpowder, . . . . .	304
Expansion, . . . . .	25	Gypsum, . . . . .	348
Explosive mixtures, . . . . .	65	Heat, . . . . .	21
Extract of lead, . . . . .	396	Heavy carburetted hydrogen gas, . . . . .	153
Febrifuge of sylvius, . . . . .	315	Hepar sulphuris, . . . . .	298
Fecula, . . . . .	207	Hombert's phosphorus, . . . . .	348
Ferment, . . . . .	163	pyrophorus, . . . . .	367
Ferrocyanate of potash, . . . . .	312	Hydracids, . . . . .	43, 262
iron, . . . . .	385	Hydrates, . . . . .	45
Fibrine, . . . . .	510	Hydriodate of potash, . . . . .	300, 315
Fire damp, . . . . .	156	Hydrocarburet of chlorine, . . . . .	255
Fixed oils, . . . . .	226, 230	Hydrogen, . . . . .	60
Flake white, . . . . .	449	its properties, . . . . .	60
Flasks, . . . . .	ix	chemical relations of, . . . . .	68
Flowers of benzoin, . . . . .	201	deutoxyde of, . . . . .	71, 75
zinc, . . . . .	408	preparation of, . . . . .	60, 64
Fluate of lime, . . . . .	350	Hydrosulphurets, . . . . .	122, 299
Fluorine, . . . . .	268	Hypo, . . . . .	45
Flux, black, . . . . .	311	Ice, . . . . .	74
white, . . . . .	311	rotten, . . . . .	74
Free caloric, . . . . .	31	Identity of light and heat, . . . . .	20
Freezing mixtures, . . . . .	338	Ignition, . . . . .	272
tables of, . . . . .	543, 544	Incandescence, . . . . .	272
Fulminating gold, . . . . .	502	Incompatibles, . . . . .	97
mercury, . . . . .	491	Indigo, . . . . .	527
platina, . . . . .	504	Ink, . . . . .	384
silver, . . . . .	497	marking, . . . . .	498
Funnels, . . . . .	viii	indelible, . . . . .	498
Furnace, . . . . .	ix	Iodate of potash, . . . . .	315
Fusible metal, . . . . .	508	Iodide of nitrogen, . . . . .	265

	Page		Page
Iodide of lead, . . . .	392	Magnesium, . . . .	356
starch, . . . .	263	Malachite, . . . .	402
sulphur, . . . .	266	Manganese, . . . .	449
potassium, . . . .	300	its compounds, 450,	456
Iodine, . . . .	227	Manganesite of potash, . .	452
preparation of, . . . .	258	Marking ink, . . . .	498
properties of, . . . .	259	Massicot, . . . .	389
tests of, . . . .	260	Measures and weights, . .	538
medical uses, . . . .	261	Medical uses of heat, . . .	40
Ioduretted hydriodate of potash, .	317	Mercurial trough, . . . .	106
Iridium, . . . .	506	Mercury and its compounds, 467,	493
Iron and its compounds, . . . .	372	tests of, . . . .	475
its oxydes, . . . .	373	Metals, . . . .	276
tests of, . . . .	374	Microcosmic salt, . . . .	340
tinned, . . . .	509	Mineral, chamcleon, . . . .	452
Ittria, . . . .	372	waters, . . . .	519
Ittrium, . . . .	372	Minium, . . . .	391
Ivory black, . . . .	142	Molybdenum, . . . .	464
Kelp, . . . .	318	Mordant, . . . .	528
Kermes mineral, . . . .	414	Morphia, . . . .	212
Keyser's pills, . . . .	491	salts of, . . . .	212
Labarraque's disinfecting liquor, .	322	Mosaic gold, . . . .	445
Lac sulphuris, . . . .	124	Mucilage, . . . .	204
Lakes, . . . .	527	Multiple proportions, . . .	9, 10
Lamp black, . . . .	142	Naphtha, . . . .	159
Lamp, safety, . . . .	157	Naphthaline, . . . .	160
Lamp, spirit, . . . .	166	Narcotine, . . . .	218
Lapis calaminaris, . . . .	410	Neutral, . . . .	45
Latent heat, . . . .	30	Nickel and its compounds, . .	505
Laws of combination, . . . .	9, 10	Night lamp, . . . .	67
Lead and its compounds, 387,	397	Nihil album, . . . .	408
its action on water, . . . .	387	Nitrate of potash, . . . .	301
oxydes of, . . . .	389, 391	Nitrite of do. . . .	305
Light, . . . .	19	Nitre, . . . .	301
Light carburetted hydrogen, . . .	156	Nitric ether, . . . .	172
its properties, . . . .	157	preparation of, . . . .	173
Liquid magnesia, . . . .	361	spirit of, . . . .	174
Litmus, . . . .	526	its medical uses, . . . .	175
paper, . . . .	526	Nitric oxyde, . . . .	84
Lime and its compounds, 343,	351	chemical relations, . . .	86, 87
cream of, . . . .	343	preparation of, . . . .	85
tests of, . . . .	344, 345	Nitrogen, . . . .	76
slaked, . . . .	342	bi-carburet of, . . . .	176
water, . . . .	343	chloride of, . . . .	253
Litharge, . . . .	390	deutoxyde of, . . . .	84
Lithia, . . . .	341	iodide of, . . . .	265
Lithium, . . . .	341	preparation of, . . . .	77, 78
Lunar caustic, . . . .	497	protoxyde of, . . . .	80, 81
Lupulinc, . . . .	226	properties of, . . . .	79
Lutes and cements, . . . .	529	Nitrous selenite, . . . .	349
Magnesia, . . . .	357	Nitrous oxyde, . . . .	82
its compounds, 358,	361	its effects on man, . . .	83
tests of, . . . .	357	inhalation of, . . . .	83
calcined, . . . .	357	preparation of, . . . .	78



	Page		Page
Nitrum flammans, . . . . .	338	Pinchbeck, . . . . .	507
Nomenclature, . . . . .	41	Platina, . . . . .	503
Nooth's apparatus, . . . . .	149	its compounds, . . . . .	503, 505
Oil, essential, . . . . .	226, 230	sponge, . . . . .	503
expressed, . . . . .	226, 230	Plaster of Paris, . . . . .	348
fixed, . . . . .	226, 230	Plasters, . . . . .	228
of vitriol, . . . . .	109	Plumbers' solder, . . . . .	508
of lime, . . . . .	348	Pneumatic trough, . . . . .	viii
of wine, . . . . .	169	Pompholix, . . . . .	407
volatile, . . . . .	226, 230	Port-fire, . . . . .	304
Olefiant gas, . . . . .	153	Potash, . . . . .	294, 295, 307
how made, . . . . .	153	tests of, . . . . .	296
characters of, . . . . .	154, 155	uses in medicine, . . . . .	296
Opium, its active principles, . . . . .	212, 217	silicated, . . . . .	371
antidotes, . . . . .	210	Potassium, . . . . .	287
detection of, . . . . .	217	its characters, . . . . .	291
Orpiment, . . . . .	440	oxydes, . . . . .	292, 297
Osmazome, . . . . .	513	preparation of, . . . . .	288, 290
Osmium, . . . . .	506	Potassiuretted hydrogen, . . . . .	297
Ox-acids, . . . . .	43	Potato, starch from, . . . . .	207
Oxy-chlorate of potash, . . . . .	315	Powder, blue, . . . . .	462
Oxygenated water, . . . . .	75	Precipitate per se, . . . . .	474
Oxy-hydrogen blowpipe, . . . . .	69	Predisposing affinity, . . . . .	18
Oxy-muriate of mercury, . . . . .	481	Prepared chalk, . . . . .	349
potash, . . . . .	313	Proof spirit, . . . . .	165
lime, . . . . .	350	Prussian blue, . . . . .	385, 386
Palladium, . . . . .	506	Prussiate of potash, . . . . .	312
Parting, . . . . .	501	Prussine gas, . . . . .	176
Patent yellow, . . . . .	391	Pulvis algarothi, . . . . .	415
Pearlash, . . . . .	307	antimonialis, . . . . .	418
Pearl white, . . . . .	449	Pyroacetic spirit, . . . . .	175
Perchlorate of potash, . . . . .	315	Pyrometers, . . . . .	29
Pewter, . . . . .	508	Pyrophorus, . . . . .	367
Philosophers' wool, . . . . .	408	Pyroxilic spirit, . . . . .	175
Phlogiston, . . . . .	60	Quadro carburetted hydrogen, . . . . .	158
Phosgene gas, . . . . .	255	Quartation, . . . . .	501
Phosphoric match bottle, . . . . .	140	Quaternary, . . . . .	45
Phosphorus, . . . . .	125	Quicksilver, . . . . .	467
antidotes, . . . . .	132	Quina, . . . . .	219
care in the use of, . . . . .	129	preparation of, . . . . .	220
its compounds, . . . . .	131	salts of, . . . . .	220
medical use of, . . . . .	132	adulterations, . . . . .	222
poisoning by, . . . . .	132	Quinine, . . . . .	219, 220, 222
preparation of, . . . . .	126, 128	Radiation of caloric, . . . . .	22
properties of, . . . . .	129	Ratsbane, . . . . .	439
the burns from, . . . . .	132	Reflection of caloric, . . . . .	23
Phosphuret of sulphur, . . . . .	140	Realgar, . . . . .	440
of potassium, . . . . .	300	Red precipitate, . . . . .	474
of calcium, . . . . .	346	sand, . . . . .	361
of lime, . . . . .	346	lead, . . . . .	391
of iron, . . . . .	375	Rhodium, . . . . .	506
Phosphuretted hydrogen, . . . . .	135	Resin, . . . . .	230
preparation of, . . . . .	136	of copper, . . . . .	400
properties of, . . . . .	138	Rochelle salt, . . . . .	328
Pictet's reflectors, . . . . .	24		

	Page		Page
Rock alum, . . . . .	365	Spirit of salt, . . . . .	248
Rust of brass, . . . . .	405	of wine, . . . . .	160
Safety lamp, . . . . .	157, 158	proof, . . . . .	165
Sal ammoniac, . . . . .	386	of sulphuric ether, . . . . .	172
alembroth, . . . . .	492	of hartshorn, . . . . .	331
catharticus amarus, . . . . .	359	of nitre, . . . . .	90
deduobus, . . . . .	305	volatilic aromatic, . . . . .	322
diureticus, . . . . .	309	mindererus, . . . . .	339
cnixum, . . . . .	306	Spongy platina, . . . . .	503
essentiale vini, . . . . .	309	inflames hydrogen, . . . . .	66
gem, . . . . .	323	Standard silver, . . . . .	508
martis, . . . . .	378	Starch, . . . . .	207
perlatus, . . . . .	328	mucilage of, . . . . .	208
polychrest, . . . . .	305	Stearine, . . . . .	227
prunelle, . . . . .	304	Steel, . . . . .	376
rupellensis, . . . . .	328	cast, . . . . .	376
rochelle, . . . . .	328	Stibiated tartar, . . . . .	417
sennertii, . . . . .	309	Strontian, . . . . .	355
Saline waters, . . . . .	523	Strontium, . . . . .	351
Salt, definition of, . . . . .	301	Strychnia, . . . . .	223
Salt of tartar, . . . . .	306	Substantive colors, . . . . .	528
wormwood, . . . . .	306	Succinates, . . . . .	203, 329
Salts of lead, . . . . .	392	Sugar, . . . . .	206
Seidlitz powders, . . . . .	328	of lead, . . . . .	394
Selenite, . . . . .	348	Sulphas potassæ cum sulphure, . . . . .	304
Selenium, . . . . .	124	Sulphates, . . . . .	108, 401, 402
Seleniuretted hydrogen, . . . . .	125	Sulphites, . . . . .	108
Sesqui, . . . . .	45, 335	Sulphur, . . . . .	103
Signal lights, . . . . .	304	flowers of, . . . . .	104
Silica, . . . . .	370	milk of, . . . . .	104
Silicated potash, . . . . .	371	precipitated, . . . . .	104
Silicum, . . . . .	369	sublimed, . . . . .	104
Silicates, . . . . .	371	Sulphuret of potash, . . . . .	298
Silicious waters, . . . . .	525	iron, . . . . .	375
Silver and its compounds, . . . . .	493	Sulphuretted hydrogen, . . . . .	116
fulminating, . . . . .	497	chemical relations, . . . . .	121
tests of, . . . . .	497	effects on man, . . . . .	122
Simple decomposition, . . . . .	14	how made, . . . . .	117, 118
Smalt, . . . . .	462	properties, . . . . .	119, 120
Smelling salts, . . . . .	335	used as a test, . . . . .	121
Soaps, . . . . .	228	tests of, . . . . .	122
Soda, . . . . .	318	Sulphuret of phosphorus, . . . . .	140
salts of, . . . . .	319, 330	of potassium, . . . . .	298
chloride of, . . . . .	321	Sulphuretted hydrosulphurets, . . . . .	123, 299
Sodaic powders, . . . . .	326	Sulphuric ether, . . . . .	167, 168
Sodium, . . . . .	317	its uses, . . . . .	170
its compounds, . . . . .	318, 330	impurities of, . . . . .	171
Soluble tartar, . . . . .	311	its gas, . . . . .	172
Sources of heat, . . . . .	37	Sulphurous waters, . . . . .	523
Spanish white, . . . . .	449	Supracetate of lead, . . . . .	394
Specific gravity, . . . . .	46, 49	Sympathetic inks, . . . . .	449, 463
modes of taking it, . . . . .	47, 48, 49	Syringe, condensing, . . . . .	38
heat, . . . . .	34	Table of vegetable acids, . . . . .	190
Speculum metal, . . . . .	507	Tannin, . . . . .	209
Splter, . . . . .	407		

	Page		Page
Tannin, artificial, . . . . .	210	Vitriol, blue, . . . . .	401
Tartar emetic, . . . . .	423	green, . . . . .	380
Tartar, . . . . .	309	white, . . . . .	410
Tasteless purging salt, . . . . .	328	oil of, . . . . .	109
Ternary, . . . . .	45	Vitriolated tartar, . . . . .	305
Terra foliata tartari, . . . . .	309	Volatile alkali, . . . . .	331
Test-papers, . . . . .	525	Water, . . . . .	71
Thermometers, . . . . .	26	changed to steam, . . . . .	74
tables of, . . . . .	542	distilled, . . . . .	72
Thorina, . . . . .	372	mineral, . . . . .	519
Thorinum, . . . . .	372	power of absorbing gases, . . . . .	73
Tin and its compounds, . . . . .	443, 446	Weights and measures, . . . . .	538
tests of, . . . . .	446	White copperas, . . . . .	410
Titanium, . . . . .	466	flux, . . . . .	311
Triacetate of lead, . . . . .	396	lead, . . . . .	394
Trona, . . . . .	327	muriate of copper, . . . . .	400
Trough, mercurial, . . . . .	106	pearl, . . . . .	449
pneumatic, . . . . .	viii	precipitate, . . . . .	492
Tungsten, . . . . .	463	vitriol, . . . . .	410
Turneric, . . . . .	526	sand, . . . . .	361
Turnsol, . . . . .	526	Wine, oil of, . . . . .	169
Turbith mineral, . . . . .	479	spirit of, . . . . .	160
Type metal, . . . . .	508	Woulfe's apparatus, . . . . .	ix
Tutty, . . . . .	408	Xanthic oxyde, . . . . .	516
Uranium, . . . . .	465	Xanthogen, . . . . .	232
Uric acid calculus, . . . . .	514	Yeast, . . . . .	163
Urinary concretions, . . . . .	514, 516	Yellow, chrome, . . . . .	461
Varnishes, . . . . .	530	dyes, . . . . .	528
Vegetable alkalis, . . . . .	210	patent, . . . . .	391
acids, . . . . .	190	Yttrium, . . . . .	372
Vegetables, analysis of, . . . . .	517	Zaffre, . . . . .	461
Veratrine, . . . . .	225	Zymome, . . . . .	209
Verdigris, . . . . .	404	Zinc, . . . . .	407
Verditer, . . . . .	403	its compounds, . . . . .	407, 411
Vermillion, . . . . .	488	sheet, . . . . .	407
Vinegar, . . . . .	191	granulated, . . . . .	407
radical, . . . . .	191	Zirconia, . . . . .	372
Vinous fermentation, . . . . .	191	Zirconium, . . . . .	372



Med. Hist.

WZ

270

M682e

1832

(1)



